Laboratory Manual of Chemistry

Armstrong and Norton

LABORATORY MANUAL

OF

CHEMISTRY

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PREFACE.

IN preparing this brief course, the authors have endeavored to select such work as will best develop the true spirit of investigation. It is not intended to present an exhaustive treatise, but rather a few of the stepping-stones over which the student of this department of nature must travel. We most thoroughly believe that the experiments selected should, as a rule, be of such a nature that the pupil can perform them for himself. The teacher should be but the guide that points out the right path, calling attention to the by-paths of error.

The work here laid out should be accompanied by some text-book. The authors have had in mind, while preparing this work, the "Manual of Chemistry," by Eliot and Storer; although it could be used with any other good text-book of elementary chemistry. It is our plan to require three hours' work per week in the laboratory and two in the recitation-room. The experiments here given, with the necessary practice in determining unknown substances by the Key, will probably occupy the class for forty weeks. Some experiments should be indicated by the teacher for omission, if the time alloted for chemistry is too short.

We believe the best results will be attained by requiring the pupil to write his notes in the laboratory at the time the experiment is performed, and not after he has been assisted to draw "bookish" conclusions from various sources. If the work is not so neatly done, or if wrong conclusions are reached, the mistakes are his own and can be corrected by pointing out, not the mistaken inference but the errors in conditions, accidents and coincidences. Let the pupil perform the experiment again in the teacher's presence, or perform it by direction under new conditions, and so finally be led to detect the error of judgment. Long descriptions of experiments are usually unprofitable to

both pupil and teacher. A simple statement of the results obtained, containing the answers to the questions asked, is usually all that should be given, unless otherwise directed. When the teacher performs an experiment, a full account of the process should be written. The "Laboratory Manual" should not be taken from the laboratory, as it offers opportunity to copy work that should be done experimentally.

We wish to acknowledge our obligations to Professor A. V. E. Young, Northwestern University, Evanston, Ill., and to Professor Le Roy C. Cooley, Vassar College, for special experiments noted in the text; to Professor F. Sanford, Leland Stanford University, California, Professor M. Delafontain, Mr. F. L. Morse and Mr. A. L. Smith, of the Chicago High Schools, for valuable suggestions in preparing the work; and also to Mr. W. E. Danner, of the Nutriment Company, Union Stock Yards, Chicago, for materials and suggestions in preparing the experiments upon Digestive Ferments.

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INTRODUCTORY WORK.

PHYSICAL PROPERTIES OF MATTER.

Experiment 1.

Materials and Apparatus.—Pieces of wood, iron, lead, wax, putty, glass, sulphur, and sugar.

Examine each of the above, and compare with water and air.

Compare with reference to color, odor, hardness, brittleness, elasticity, and form.

Do they occupy space?

Have they weight?

Do they appeal to each of the senses? Define matter

Experiment 2.—Specific Volume.

Materials and Apparatus.—Same as in Experiment 1 and metric rule, balances, pocket-knife, test-tube, and burette.

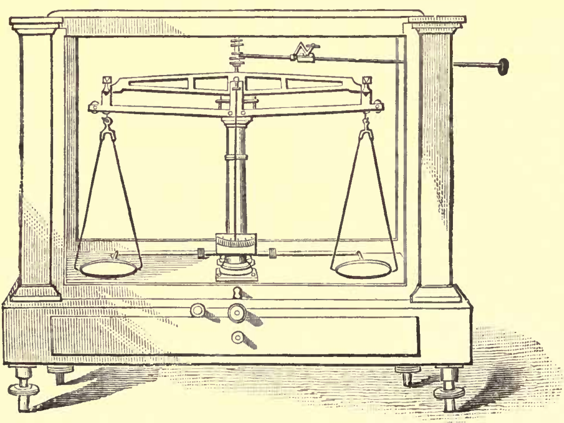


Fig. 1.—Chemical Balance.

(*a*) Allow 5 or 10 c. c. of water to run into a test-tube from the burette. Note exact point at which water stands in your tube.

Pour back the contents of your test-tube into the burette.

Does the water in the burette stand where it did at the beginning of the experiment? Explain.

Carefully mark off on the side of one of your test-tubes the point to which 5 c. c. water rises. Also, if convenient, mark 2.5 c. c. and 10 c. c.

N.B.—In all subsequent experiments, when you are requested to take 5 c. c. or 10 c. c. of any liquid, measure out the amount as accurately as possible, using your graduated test-tube.

Measure out 25 c. c. of water in a beaker by filling your graduated test-tube five times to the 5 c. c. mark. Now pour this water into the burette and note how much you must allow for errors in your work when you measure out large quantities of water in this manner.

(*b*) Make cubes of wax, putty, lead, and wood, each weighing a gram. Measure carefully with metric rule and compare volumes, using c. c. for the unit.

What is the volume of 1 gram of water at the temperature of its greatest density?

Weigh 5 c. c. of water in test-tube; empty and weigh tube. What result do you obtain as an answer to the preceding question?

Now compare volumes of above substances to that of 1 gram of water, and state their specific volume.

Experiment 3.—Specific Gravity.

Materials and Apparatus.—Balances, test-tube (rather large), sulphur, or some other substance of which to determine the specific gravity.

Find specific gravity of the substance given you, in any way you choose.

One method.—Carefully weigh one of the cubes (*b*, Exp. 2). Fill a beaker so full of water that nothing can be added without causing an overflow, and place the beaker in a porcelain dish.

Now immerse the cube in the water in the beaker and collect the overflow in the porcelain dish and weigh.

What was the exact weight of the cube?

What is the volume of the water which overflows?

What is the weight of this volume of water?

How many times the weight of the water is the weight of the solid? This is the "specific gravity" of the solid.

Another method.—Weigh 10 or 15 grams of dry sulphur in lumps. Weigh test-tube filled with water and corked. Put in sulphur, fill with water, and weigh again.

Specific gravity = weight of sulphur ÷ weight of equal volume of water.

CHANGE.—Performing an experiment in physics or chemistry consists in surrounding some form of matter with certain conditions, which should be carefully noted, and in watching the result. Accounting for the changes which have taken place comes next, and is often the most difficult portion of the work.

Experiment 4.—Some Changes Due to Heat.

(*a*) Fill test-tube with water so that it stands up above the edge of the tube. Warm gently. What happens? How do you account for it? What change takes place in the water? Does this change involve a change in the *nature* of the substance?

(*b*) Fit a short piece of glass tubing into a cork; fill test-tube with water and insert cork, forcing the liquid up to a point where it can be seen in the tube. Warm the water in the test-tube as in (*a*), and answer the same questions.

(*c*) Perform the same experiment, having air in the test-tube and a drop of water in the glass tubing. Answer same questions as in (*a*). This is an air thermometer. What is a differential thermometer?

Experiment 5.—Latent Heat.

Apparatus.—Ring-stand, evaporating dish or glass beaker, and chemical thermometer. Take a quantity of pounded ice or snow—enough to fill your beaker. Place on wire-gauze or sand-bath, on your ring-stand, and heat.

Note temperature when melting begins.

Note temperature while melting proceeds.

Note temperature just as last particle is melted.

How much time was required? State as exactly as possible.

Continue to heat to boiling point. Note temperature at which it begins to boil.

Note time required to heat the water—after the ice was melted—to the point where it began to boil.

Boil water all away and note time required.

Compare the three results obtained.

How do your results differ from figures given in the text?

What is latent heat?

State all the changes which take place in above experiment.

Is the substance changed? To what extent?

Are ice, water, and steam one substance?

Experiment 6.—Changes Affecting the Boiling Point.

(*a*) Find temperature of boiling for some other liquid.

(*b*) Make strong solution of salt in water and note how boiling point is changed.

N.B.—Care should be taken in the following to avoid explosion. Note these points of caution:

Just as soon as boiling begins, remove heat and put in stopper. In inverting tube be careful not to burn your hands. Now pour on a quantity of cold water, and as soon as water in tube begins to boil for the first time, so that you are able to see the phenomenon, bring your experiment to an end and do not try to make the water boil a second time.

(*c*) Boil water in test-tube (half full), insert stopper, and invert.

Remove the heat and bring the water to boil again by pouring on cold water.

How does the cold water affect the matter above the water?

State your conclusion in regard to (*c*).

Experiment 7.—Evaporation, Vaporization, Distillation.

Boil water and collect steam on piece of cold glass or porcelain.

What changes take place?

What is distillation?

Place a few crystals of iodine in test-tube, heat, and describe changes which take place.

Examine substance formed on the sides of the tube near the top, where it is cool.

Does it resemble the substance with which you started?

What is sublimation?

To what extent have the substances been changed in this experiment?

QUESTIONS ON EXPERIMENTS 4, 5, 6, AND 7.

In any of the changes produced, has the nature of the substance been altered?

Put a few drops of ether on your wrist and blow on it; account for the sensation produced.

What does this suggest with regard to perspiration?

Does melting take place through a wide range of temperature?

Does vaporization?

Which form?

Experiment 8.—Chemical Change.

(*a*) Try to dissolve a small quantity of flowers of sulphur in bisulphide of carbon.

Try to pick up small particles of iron with a magnet, using iron filings or powdered iron. This shows one of the properties of iron by which you can recognize it in the future.

Try some dilute hydrochloric acid on some sulphur. Is any gas given off having an odor?

Try same acid on particles of iron and answer same question as to gas in this case.

(*b*) Mix thoroughly small quantities of powdered iron and flowers of sulphur.

Is a new substance formed?

Is the sulphur still sulphur?

Try bisulphide of carbon.

Is the iron still iron?

Try magnet.

(*c*) Weigh out 3.2 grams sulphur flowers, also 5.6 grams powdered iron. Mix thoroughly and place in hard glass test-tube and heat over burner until the mass suddenly glows vividly.

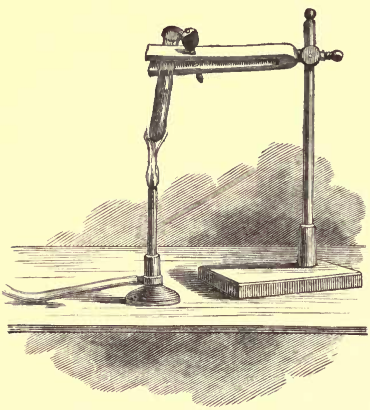


Fig. 2.—Union by Heat.

Set aside to cool. When it has cooled, examine for sulphur and for iron.

Is the iron still iron?

Is the sulphur still sulphur?

Try substance with hydrochloric acid and note odor of gas given off and its effect upon paper moistened in solution of lead acetate.

How many substances had you to begin with?

How many have you now?

How does this change differ from those changes noted in all previous experiments?

With what kind of changes is chemistry concerned?

Note.—Whenever an experiment is marked "For the Teacher," it should be performed by the teacher in the presence of the class.

Experiment 9.—Chemical Change, Continued.

(*For the Teacher.*)—With the apparatus shown in Fig. 3, the teacher is to perform the familiar experiment of electrolysis of water, while the pupils take notes of all that goes on and answer on opposite page any questions which the teacher may emphasize.

Note following points:

How many substances have you as a result of your experiment? How many had you to begin with? How does experiment in this respect differ from preceding? What is analysis? What is synthesis?

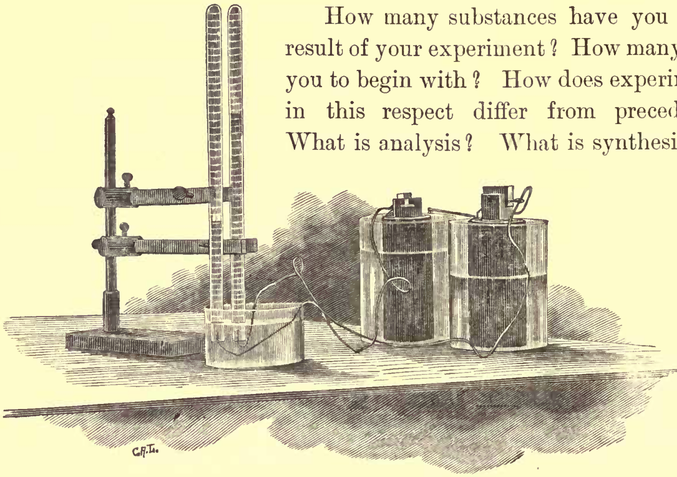


Fig. 3.—Electrolysis of Water.

Note.—The above experiment is here introduced merely to show the difference between chemical and physical changes. The analysis of water will be discussed later.

Experiment 10.—Solution.

Materials and Apparatus.—Large beakers, sugar, filter-paper, funnel, funnel-stand.

Take 50 c. c. of water in a beaker and add to it 10 grams of finely pulverized sugar, stirring with a glass rod. Does the sugar disappear? Where has it gone? Is this a physical or a chemical change? Has the sugar lost its essential characteristics? Taste it. Fold a filter-paper and place it in the glass funnel. Support this on the funnel-stand and pour the solution into it. Does the sugar pass through as readily as water, or can you separate out the sugar? Can you separate them by boiling away the water? Evaporate until you get a thick sirup, and then set aside for several days. Can you dissolve clay in water? Try it. Pass it several times through two thicknesses of filter-paper. Is this a solution? What is a solution?

Experiment 11.—Hot Solution.

Materials and Apparatus.—Beaker, alum, mortar, and pestle.

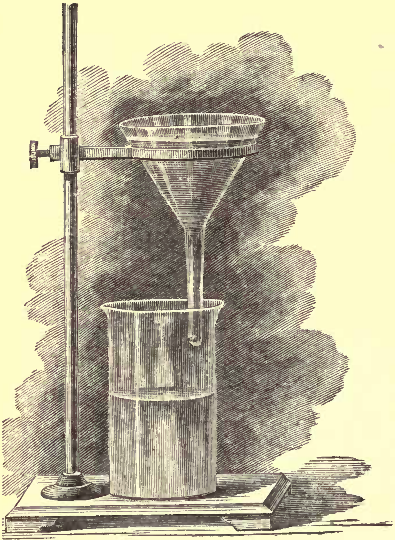


Fig. 4.—Filtration.

Take 25 grams of cold water in a beaker and add to it 10 grams of pulverized alum. Stir for some time. Does all the alum dissolve? Why will it not dissolve any amount you may add? A sponge full of water is said to be saturated, so this is called a saturated solution. Now heat the solution by placing the beaker upon the sand-bath over the gas-burner. Does the alum all disappear? What do you infer as to the effect of temperature upon the amount of alum held in solution? Save the alum solution 24 hours or more for another experiment.

Experiment 12.—Crystallization.

Materials.—Sugar solution and alum solution of last experiments.

Examine the two solutions set aside. What has occurred? Has any of the water evaporated? Examine some of the most regular crystals of each. Have they definite forms? Describe the most common of each. If you had sealed up the solution of sugar, would crystals have formed? If you had sealed up a hot solution of alum, would you expect to find crystals? Try these experiments, being careful that the solutions are kept air-tight.

Experiment 13.—Latent Heat of Solution.

Materials and Apparatus.—Test-tubes, basin, thermometer, ice, salt.

Mix equal parts of snow, or ice, and salt in a basin. Test the temperature from time to time with a thermometer as the mass is stirred. Does the act of going into solution consume heat? If a stirring-rod is made by putting a small test-tube inside a larger one, with the space between filled with water, the water will become frozen. The heat consumed in changing the solids to solution is *latent* or *insensible* heat. If crystallization is the opposite of solution, should we not expect that latent heat is given off? Is it warmer just before a snow-storm?

Experiment 14.—Freezing Mixture. Latent Heat, Continued.

(*For the Teacher.*)—Take 9 parts by weight of phosphate of sodium and 4 parts by weight of dilute nitric acid. Cool down the ingredients by snow and salt to near 0°, then mix the ingredients and test the temperature with a thermometer.

How low a temperature have you reached? How is ice made artificially? In what experiment have you made heat latent by evaporation? Blow upon the hand with the mouth wide open. Do so again with the lips nearly closed. In which case did the breath seem cold? Why was it cold? Does condensation produce heat?

Experiment 15.—Allotropism.

Materials and Apparatus.—Test-tube, beaker, sulphur, gas-lamp.

Put 5 grams of sulphur into a test-tube and melt slowly over the gas-lamp until it begins to boil. Do not hold the mouth of the test-tube too near the flame for fear the vapor may take fire. Notice changes of color as the sulphur melts. Does it become thin enough to flow? Does it when about to boil? Pour a fine stream into cold water, and allow same to cool slowly. Are these two parts alike? If not, how do they differ? Examine the soft sulphur after two or three days. Has it become hard and yellow?

Fig. 5.

When a substance assumes two or more forms that differ only in physical properties, it is called *Allotropic.* The diamond, charcoal, coke, and lamp-black are allotropic forms of carbon.

NOTE.—Where among these changes lies the domain of chemistry? All the changes we have made suggest two distinct groups of changes, with one intermediate group.

1st, Those changes that do not affect the nature of the substance; 2d, Those changes which do affect the nature of the substance; and 3d, Changes that seem to have affected the nature of the substance, but in reality have only changed the physical state.

Chemistry is concerned directly with the latter two, though indirectly with the first-mentioned class of changes.

REVIEW QUESTIONS.

What is a physical change? A chemical? A physico-chemical? Which of the experiments you have tried belong to physical? Which illustrate chemical and which physico-chemical? Give other illustrations of each.

Experiment 16.—Solution assists Chemical Change.

Materials and Apparatus.—Mortar and pestle, test-tubes, bicarbonate of soda, tartaric acid.

Mix in a dry mortar half a gram each of dry tartaric acid and bicarbonate of soda (cream of tartar and baking-soda). Does a chemical change occur? Now make solutions in test-tubes of each of the same ingredients and pour into one test-tube. Does a chemical change occur? Hold the thumb over the mouth of the test-tube for a moment. Is some gas being expelled? Why is there chemical change in one case and not in the other? Pour water into the mortar. What occurs? How does solution help chemical change?

Experiment 17.—Divisibility of Matter.

Materials and Apparatus.—Purple aniline, a litre or quart flask.

Into a litre (1000 c. c.) of water in a flask put one decigram of aniline powder. Does it dissolve? Pour out half of the water after it is thoroughly mixed and add as much clean water. How much aniline have you now? Pour out half and fill as before. How much aniline left? Repeat this as long as you can see a trace of aniline in the water, keeping account all the time of the amount of aniline remaining in the flask. Divide by 1000, since you have 1000 c. c. of water. How much aniline have you in each c. c.? Imagine this division carried on until you have reached the smallest possible particle of aniline. Call it a *Molecule.* Thus each substance may be conceived to be made up of molecules. In how many ways can molecules differ?

Experiment 18.—Porosity of Matter.

Apparatus.—Beakers, air-pump.

Fill a glass with cold water and stand it in a warm room for half an hour. Do you notice air-bubbles rising? Why does air come from the water under these circumstances? Put a fresh glass of water under the receiver of the air-pump and exhaust the air. Can you pump air out of water? How can a vessel full of water contain air? When the air is exhausted is the volume of water less? When alum went into solution did it displace part of the water, that is, increase its volume? What must be the shape of molecules of water to allow small particles of a solid or gas to fill part of the same space? Could you illustrate with oranges and peas? Will a c. c. of water saturated with alum, hold as much sugar in solution as a c. c. of water containing no alum? Try it.

Experiment 19.—Atoms.

Materials and Apparatus.—Sugar, platinum-foil, gas-burner, tongs.

Weigh out a decigram of sugar, place upon the platinum-foil, and burn it over the flame.

Is the charred mass lighter than a decigram? What is it? What has become of the rest of the substance?

Is this a chemical or a physical change?

If you had been able to take a molecule of sugar, and had thus burned away part of it, would you then have less than a molecule of matter?

Would it be sugar? How do you know?

What force divided the molecules of sugar into parts of different kinds?

Call these parts of a molecule *Atoms.*

Experiment 20.—Compounds and Elements.

(*For the Teacher*) *Materials and Apparatus.*—Chemical battery U-tube, sulphuric acid.

Decompose water and develop the answers to the following questions with others that may arise:

How are the volumes of the two gases that arise from the poles of the battery related? Is this a chemical change?

Is the same kind of gas found in each tube?

Test by a burning splinter. Do they behave alike?

What must be the proportion of each of these gases in each molecule of water?

If a substance is composed of two or more simple substances, the first is *compound* with reference to the constituents. Hence we call a substance that the chemist is unable to divide into two or more substances an *element*, and two or more elements combined, a *compound.*

Where would you class each of the following:—sugar, paper, wood, sulphur, iron?

NOTE.—The table on page 28 gives a list of the most common elements that form the common compounds, though some new process may be discovered by which some of these may be divided into two or more other elements. New elements may be discovered from time to time, though not probably of great practical importance.

The symbols are used for convenience, and should be learned as fast as the elements are studied.

Atomic weights are simply for reference. Symbols marked thus (′) are called monads; (″), diads; (‴), triads; (⁗), tetrads. (For explanation of these terms, see Exp. 39.)

SYMBOLS AND ATOMIC WEIGHTS OF THE COMMON ELEMENTS.

ELEMENT.	SYMBOL.	ATOMIC WEIGHT.	ELEMENT.	SYMBOL.	ATOMIC WEIGHT.
Aluminum	Al''''	27.04	Lead (Plumbum)	Pb''	206.4
Antimony (Stibium).	Sb'''	119.6	Magnesium	Mg''	23.94
Arsenic.	As'''	74.9	Manganese	Mn''	54.8
Barium	Ba''	136.9	Mercury (Hydrargyrum)	Hg''	199.8
Bismuth	Bi'''	207.3	Molibdenum.	Mo''	95.9
Boron	B'''	10.9	Nickel (?)	Ni''	58.56
Bromine	Br'	79.76	Nitrogen	N'''	14.01
Cadmium	Cd''	111.7	Oxygen	O''	15.96
Calcium	Ca''	39.91	Phosphorus	P'''	30.96
Carbon	C''''	11.97	Platinum.	Pt''''	194.3
Chlorine	Cl'	35.37	Potassium (Kalium)	K'	39.03
Chromium	Cr''''	52.45	Silicon	Si''''	28.
Cobalt (?)	Co''	58.74	Silver (Argentum)	Ag'	107.66
Copper	Cu''	63.18	Sodium (Natrium).	Na'	23.
Fluorine	F'	19.06	Strontium	Sr''	87.3
Gold (Aurum)	Au'''	196.7	Sulphur	S''	31.98
Hydrogen	H'	1.	Tin (Stannum)	$\mathrm{Sn}''\&''''$	117.4
Iodine	I'	126.54	Zinc	Zn''	65.1
Iron (Ferrum)	$\mathrm{Fe}''\&'''$	55.88			

Experiment 21.—Compounds, Continued.

How can so few elements make up the innumerable compounds that are found in nature? Take the four digits, 1, 2, 3, and 4. Arrange them in all the different ways possible to form numbers. Thus, 1234, 1324, 4321, etc. How many numbers can you form? If you arrange them in all the possible combinations of twos, threes, and fours, how many numbers can you form? If there were but four elements, and this were the only way they could unite, how many compounds would there be? How many words in the English language formed from 26 letters? Has the limit been reached?

Experiment 22.—Analysis.

Materials and Apparatus.—A hard glass test-tube, red oxide of mercury, a long splinter of pine, gas-lamp.

Put a gram of red oxide of mercury (**HgO**) into a hard glass test-tube. Heat it slowly over the gas-burner. Thrust a glowing splinter down the test-tube from time to time to determine if a gas like one of those obtained by the decomposition of water is being given off.

Is it true? Which one is it? Is there a new substance found in the test-tube? What does it look like? What force separated oxide of mercury into two parts? Is oxide of mercury a compound? Of what is it composed? This process is called *Analysis.* In what other experiment have you used this process?

Experiment 23.—Synthesis.

Apparatus.—A lamp chimney or a large glass tube, gas-burner.

Hold a cold, dry lamp-chimney over the burning gas-lamp for a moment. Is water condensed upon the inside? Where did it come from? Illuminating gas is made up largely of hydrogen. Air contains a large quantity of oxygen. When these two unite what compound do they form? In what experiment have we performed the opposite of this experiment? Combining elements to form compounds is called synthesis. In what experiment have you used this process before? What did it illustrate? What force did you employ in each?

Experiment 24.—Acids and Alkalies.

Materials.—Litmus, acids, caustic soda, ammonium hydrate.

Put a drop of hydrochloric acid upon a piece of blue litmus paper. How does it affect it? Try other acids. Do they all have the same or different effects? Put a drop of ammonium hydrate upon the spot made by the acid. Try caustic soda. Do caustic soda and ammonium hydrate behave alike? Try acid and red litmus, then try ammonium hydrate, sodium hydrate, calcium hydrate, etc., with blue litmus. In which cases do you get a decided result?

What law can you state from these experiments as to the difference between acids and hydrates? Hydrates are a type of a class of substances called *bases*. How can you tell these bases from acids?

All elements may be classified in reference to their power of forming bases or acids; some, however, may form both a base and an acid. (See list of acid-forming and base-forming elements, below.)

Put a drop of sulphuric acid in 5 or 6 c. c. of water. Taste it. Do so with acetic and hydrochloric acids. What taste have they?

What elements have all acids in common? (See Symbols.) Define an acid. Call all the compounds you find that turn red litmus blue, *Alkalies*.

Common Base-forming Elements.				Common Acid-forming Elements.	
Al	**Cr**	**Mn**	**Na**	**As**	**F**
(NH_4)*	**Co**	**Mg**	**Sr**	**B**	**I**
Sb	**Cu**	**Hg**	**Sn**	**Br**	**N**
Ba	**Au**	**Ni**	**Zn**	**C**	**O**
Bi	**H**	**Pt**		**Cl**	**P**
Cd	**Fe**	**K**		**Cr**	**S**
Ca	**Pb**	**Ag**		**Cy** †	**Si**

* This compound of nitrogen and hydrogen is classed with elements because it forms a base like sodium and potassium. The radical $\mathbf{NH_4}$ is called ammonium in distinction from the gas $\mathbf{NH_3}$ called ammonia.

† **Cy** is a compound of carbon and nitrogen, $\mathbf{C_2N_2}$, which acts like an element in forming cyanic acid. The radical is called cyanogen.

Experiment 25.—Neutralization, Formation of Salts.

Materials and Apparatus.—Beaker, evaporating dish, water-bath, sodium hydrate, hydrochloric acid, litmus.

Slowly drop **HCl** into 25 c. c. of the solution of sodium hydrate, as long as it shows chemical action. Now test with litmus and add a drop of acid or alkali until it turns the color neither red nor blue. It is now neutral, so far as you can perceive by the litmus. A change must have occurred in the chemical nature of both. If so, new substances may be looked for. Transfer to the evaporating dish and place upon the water-bath over the gas-lamp to evaporate. Taste the dry substance obtained. What is it ?

Complete the equation, $\mathbf{NaHO + HCl = NaCl} + ?$

NaCl contains the first element of the alkaline hydrate and the last element of the acid. What is the symbol for common salt ? Is this salt acid, alkaline, or neutral ? Try the following and see if you can neutralize and obtain a salt : $\mathbf{H_2SO_4 + 2NH_4HO} = ? + ?$ Evaporate as before. Is this salt like the first ? When zinc and **HCl** were used to liberate hydrogen, was a salt formed ? Define a salt.

Is zinc capable of turning red litmus blue ? Is it capable of neutralizing the power of an acid ? Drop a grain of zinc into a test-tube and cover with dilute hydrochloric acid. Is there chemical action ? Is hydrogen given off ? Allow it to stand over night. Is there any acid property left ? Is there a solid substance formed differing from zinc ? Compare with chloride of zinc. Call all elements that can neutralize an acid, a base.

Are all alkalies that you have tried capable of forming the base of a salt ? It will be noticed that when the base is a monad it combines equally with **HCl**. Thus, $\mathbf{HCl + Na = NaCl + H}$. But when a diad base combines with **HCl** we must take two molecules of the acid. Thus, $\mathbf{Zn + 2HCl = ZnCl_2 + 2H}$, since zinc is a diad and hydrogen a monad. Again, if sodium **Na** combines with $\mathbf{H_2SO_4}$ it takes two atoms of **Na** to form the neutral salt, since there are two atoms of **H**. Thus, $\mathbf{H_2SO_4 + 2Na = Na_2SO_4 + 2H}$.

It will be seen that the number of hydrogen atoms standing first in the symbol shows the combining power of the acid, and that this hydrogen is simply thrown out to give place to a base of the same number of hands or combining powers. Call the combining power of an acid its *basicity*. As there are but few common acids, it is important to know the basicity of each so you may use them readily in writing symbols.

TABLE.—BASICITY OF ACIDS.

MONOBASIC ACIDS.	DIBASIC.	TRIBASIC.	TETRABASIC.
Hydrochloric, HCl.	Sulphuric, H_2SO_4.	Phosphoric, H_3PO_4.	Hydroferrocyanic, H_4FeCy_6.
Chloric, $HClO_3$.	Sulphurous, H_2SO_3.	Arsenic, H_3AsO_4.	
Hydroiodic, HI.	Hydrosulphuric, H_2S.	Arsenious, H_3AsO_3.	
Iodic, HIO_3.	Carbonic, H_2CO_3.	Hydroferricyanic, H_3FeCy_6.	
Hydrobromic, HBr.	Chromic, H_2CrO_4.		
Nitric, HNO_3.	Oxalic, $H_2C_2O_4$.		
Nitrous, HNO_2.	Tartaric, $H_2C_4H_4O_6$.		
Acetic, $HC_2H_3O_2$.			

The above table shows the combining power or basicity of each of the most common acids. As before stated, the atom or atoms of hydrogen standing first in the symbol may be thrown out by a base to form a salt. In this light an acid may be regarded as a salt of hydrogen, and the name hydric chloride and hydric sulphate used instead of hydrochloric and sulphuric acid.

The group of atoms following hydrogen in the symbol of sulphuric acid (H_2SO_4) is called an *acid radical.* SO_4 is the acid radical which combines with a base to form a salt.

All salts in which SO_4 is found will be called a *sulphate.* Salts formed from H_2CO_3 will have the radical CO_3, and are called *carbonates.* When there are two acids formed by the same elements, differing only in the number of atoms of oxygen, as H_2SO_4 and H_2SO_3, the name of the acid containing the greater number of oxygen atoms terminates in **ic** and that of the one containing the less in **ous.**

When the name of the acid terminates in **ic** the salt formed terminates in **ate.**

When the name of the acid terminates in **ous** the salt formed terminates in **ite.**

Thus, nitr*ic* acid forms nitr*ates*, while nitr*ous* acid forms nitr*ites*. If an acid contains no oxygen, as **HCl**, its salts terminate in **ide.** Thus, from **HCl**, chlor*ides;* from **HBr**, brom*ides*, etc. Regarding an acid as a salt of hydrogen, H_2SO_4 would be called hydrogen sulphate, or sulphate of hydrogen, or hydric sulphate. Replace the hydrogen by a base, as sodium, we have

$$H_2SO_4 + 2Na = Na_2SO_4 + 2H.$$

The salt, according to the above, would be called sodium sulphate, sulphate of sodium, or sodic sulphate.

Experiment 26.—Indestructibility of Matter or Persistence of Mass.

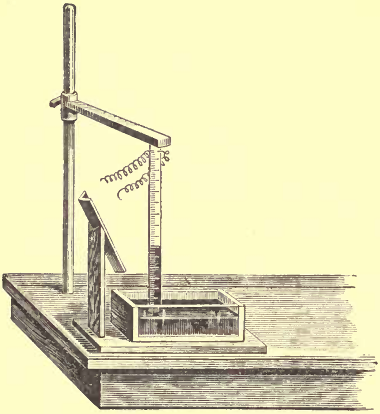


Fig. 6.—EUDIOMETER.

(*For the Teacher.*)—(*a*) Balance electrolysis apparatus on scale-pan. Decompose the water, collecting the gases. As this chemical change takes place is there a change of volume? Is there any change of mass? Is any matter destroyed?

(*b*) Perform experiment in synthesis of water in eudiometer. Have eudiometer balanced before and after passing the spark. Is there any loss of mass?

Experiment 27.*—Indestructibility of Matter, Continued.

(*For the Pupil.*)—Place less than a gram of mercuric sulphocyanate in an open dish upon the balance-pan. Counterpoise it, then ignite it with a hot iron or a match, and note effect upon the equilibrium of the balance.

Experiment 28.—Law of Definite Proportions.

(*a*) Evaporate a few drops of hydrochloric acid solution, then a few drops of ammonia solution.

(*b*) With some convenient measure (*e.g.*, a test-tube with a slip of paper gummed upon its side) measure out two equal portions of strong ammonia water. Neutralize exactly with strong hydrochloric acid solution (using litmus paper) these two portions added together in an evaporating dish, the weight of which has been determined. Note carefully the amount of acid used. Evaporate the contents of the dish to dryness on the water-bath and weigh the salt obtained.

(*c*) Take again the same quantity of hydrochloric acid and add to it only one measureful of ammonia water; evaporate as before, and weigh the salt obtained.

(*d*) Take again the same quantity of acid and add to it three measurefuls of ammonia water; evaporate as before and weigh the salt obtained. What is the ratio between the quantities of salt obtained in the three cases? What is the logic of this experiment?

* Experiments Nos. 27, 28, and 29 were prepared by Professor A. V. E. Young, Professor of Chemistry in the Northwestern University at Evanston, Ill.

Experiment 29.—Law of Multiple Proportions.

NOTE.—In the following experiment two or more pupils may work together.

(*a*) Take iodine and mercury in the proportions of 254 parts by weight of the former and 200 parts of the latter. For the actual experiment weigh out accurately (in some glass vessel) 6.3 grams of iodine. Then weigh out 5 grams of mercury and transfer it to a mortar previously weighed. Add two or three drops of alcohol and then a small portion of the iodine. Triturate thoroughly until the mass is dry. This should yield a red powder. A drop of alcohol added now should be but slightly, if at all, colored by the iodine.

If the powder is not a bright red, and the alcohol is colored by iodine, the trituration must be continued with perhaps a little warming in the water-bath. When this operation is finished, weigh mortar and contents. Then take a small portion of the red powder, such as would be taken up on the point of a penknife, transfer it to a test-tube, and add a little alcohol and boil. Be careful that the vapor of alcohol does not take fire at the gas-jet. The powder, or a portion of it, is dissolved by the boiling alcohol and redeposited, on slow cooling, in shining crystalline scales of a brilliant scarlet.

Sublime another small portion of the red powder in a dry test-tube; note the red and yellow sublimate. Rub the yellow with a glass rod; note the effect. The substance obtained in this experiment is mercuric iodide, symbol $\mathbf{HgI_2}$, soluble in hot alcohol, subliming without decomposition. The red and yellow sublimates are allotropic forms of the same composition. By gentle friction the yellow is converted into the red.

(*b*) Take 5.6 grams of the mercuric iodide obtained in (*a*) [equal to half the yield (6.3 plus 5)] and 2.5 grams of mercury [equal to half the amount used in (*a*)]. Add the mercury, in small portions at a time, to the mercuric iodide in a mortar and triturate thoroughly, using considerable pressure. This should yield a yellowish-green powder. If the color is not satisfactory, continue the trituration with the addition of a few drops of alcohol. If still unsatisfactory, allow it to stand for several hours. When finished, weigh mortar and contents. Treat small portion as in (*a*) with boiling alcohol [remember caution]. Heating another portion in dry tube turns it brick-red and yields a sublimate as in (*a*), and mercury.

The green powder thus obtained is mercurous iodide, symbol $\mathbf{HgI}$; insoluble in hot alcohol and decomposed by heat into mercury and mercuric iodide. What are the relative proportions of mercury and iodine in the two substances? What is the logic of this experiment?

Experiment 30.—Chemistry of Water.

(*For the Teacher.*)—Trough of water, metallic sodium, wire-gauze, bottle for collecting gas, piece of wire or iron tongs.

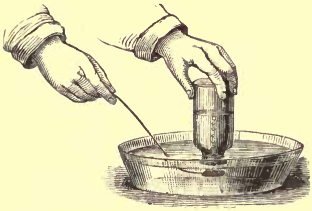


Fig. 7.

Roll up a piece of metallic sodium the size of a pea in a small piece of gauze. Invert a bottle full of water in the water-trough. By means of a stiff wire or a pair of tongs, thrust the gauze containing the sodium under the mouth of the bottle. Sodium will rapidly unite with one of the elements of the water, setting free the other. In Experiment 20 we learned that water is composed of two gases; that one burns with a blue flame, hydrogen. The other will not burn, but makes a burning splinter burn more rapidly, oxygen.

Test this gas. Which is it? When hydrogen burns what compound is formed? Write an account of this experiment and be sure to answer all of the above questions.

Experiment 31.—Distillation.

(*For the Teacher*) *Materials and Apparatus.*—Large retort or distill, condenser, platinum-foil, artesian well-water if possible, or any mineral water, gas-lamp.

Distill water before the class. What proof can you see that impurities are left in the retort? Does well-water contain solid matter in solution? Can you tell by the clearness? Put a few drops of well-water upon a clean platinum-foil. Hold it over the gas-lamp so it will not boil but evaporate rapidly. Is there a sediment left upon the foil? Try the same with distilled water. Why is distilled water valuable in making solutions? Why has it not an agreeable taste? Why are mineral waters desirable? How do they acquire these properties?

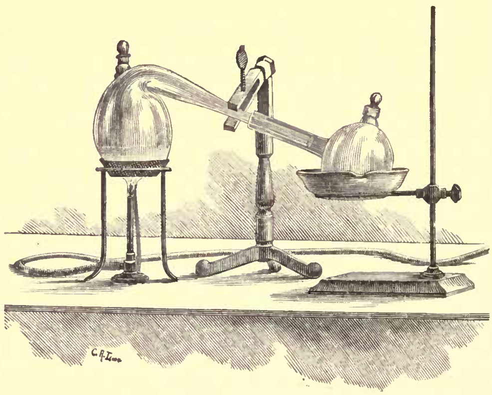


Fig. 8.—Distillation.

Experiment 32.—Water of Crystallization.

Materials and Apparatus.—Glass tube four inches long, crystals of zinc sulphate, sugar, splinters of wood.

Seal up one end of a glass tube by a blow-pipe and gas-lamp. Examine some crystals of zinc sulphate as to color, luster, etc. Put them into the test-tube and heat the lower end gently. Do the crystals change? If they contain water it will be driven out and condense near the upper end of the tube. Empty the tube when satisfied as to the presence of water; dry it, and when cool, introduce sugar, and heat as before. Does it contain water? Dry the tube, and when cool, try wood, a fiber of meat, a green stem, etc. Water in crystals is called *water of crystallization.* How can you tell the amount of water in a pound of granulated sugar? A pound of steak? How much of the human body is water?

Experiment 33.—Efflorescence and Deliquescence.

Materials and Apparatus.—Two watch-glasses, sodium sulphate, and calcium chloride.

Put some crystals of sodium sulphate upon a watch-glass and leave exposed to the air for an hour or more. Do they lose water as the zinc sulphate did by heating? Are the crystals destroyed? When crystals lose their water of crystallization by exposure to the air, they are said to effloresce. In the same way expose some calcium chloride to the air and leave over night. How has it behaved? Where did the water come from? This is said to deliquesce. How do these two salts stand with respect to their attraction for water?

If you wished to dry the moisture out of air, what could you take to do it with? What precaution is needed if you wish to keep these two salts in good condition for future use? Why?

Experiment 34.—Impurities in Drinking-water. Organic.

Materials and Apparatus.—Beakers, sulphuric acid, solution of potassium permanganate, samples of water from different sources, including one from a ditch or pool, and distilled water.

Fill the beakers, each with a sample of water, and one with ditch-water and another with distilled water. Put a few drops of sulphuric acid into each and then enough of the permanganate solution to give the water in each a deep purple tint, as nearly alike as possible. Set in a warm place for an hour. Organic matter bleaches out the color, or rather decolorizes the permanganate solution. Which one shows most and which one least organic matter? Why? How does water acquire organic matter? What is organic matter?

Experiment 35.—Impurities in Drinking-water. Chlorides or Chlorine.

Materials and Apparatus.—Beakers, silver nitrate, nitric acid, ammonium nitrate, samples of water, distilled water, and common salt.

Concentrate by boiling 50 c. c. of drinking-water to 25 c. c.; add a few drops of nitric acid and then a few drops of silver-nitrate solution. Now take 25 c. c. each of distilled-water and salt-water solution, in separate beakers, and treat in the same way with acid and silver nitrate. These two are only for comparison, as distilled water will contain no chlorine, and salt water a great deal, as salt is composed of equal volumes of sodium and chlorine (**NaCl**).

A chloride, as you see by the salt solution, becomes milky by the addition of silver nitrate. Does the sample of drinking-water from the pond? Does that from a deep well? Distilled water? Sewage is a source of chlorine in drinking-water, and such water should be avoided, unless it comes from the presence of common salt, as in salt springs, in which case it does no harm.

Experiment 36.—Hydrogen.

Materials and Apparatus.—Test-tube, granulated zinc, hydrochloric acid.

Put a few grains of zinc into a test-tube and cover with dilute hydrochloric acid. What occurs? Is it a chemical change? After the action has continued for a few minutes, apply a lighted splinter to the mouth of the test-tube. What occurs? Does this prove that chemical change has occurred? Such bubbling is called *effervescence.* What is the gas? In what experiments have you made or seen it made before? Try zinc and dilute sulphuric acid. Is the same gas given off? Again try iron and sulphuric acid. Do you find the same gas? Since zinc and iron are both elements, could the hydrogen have come from them? Look at the symbol on each of the acid bottles, **HCl** and $\mathbf{H_2SO_4}$. What element in common? To express this in chemical language, we write: $\mathbf{2HCl + Zn = ZnCl_2 + 2H}$. This reads, two molecules of hydrochloric acid and one atom of zinc form one molecule of zinc chloride and two atoms of hydrogen.

Again, $\mathbf{H_2SO_4 + Zn = ZnSO_4 + 2H}$ reads, one molecule of sulphuric acid and one atom of zinc form one molecule of zinc sulphate and two atoms of hydrogen. Evaporate the latter nearly to dryness; set aside for twenty-four hours, and see if you have crystals of a substance that is neither zinc nor sulphuric acid. Compare them with zinc sulphate.

NOTE.—A number before a molecule multiplies the molecule; hence, multiplies each of the elements in the molecule. Thus, 2**HCl** means two atoms of hydrogen and two of chlorine. A number placed below and after a letter multiplies that element only. Thus, $ZnCl_2$ means that the molecule of zinc sulphate contains one atom of zinc and two atoms of chlorine. What symbol expresses a molecule of water?

Experiment 37.—Hydrogen, Continued.

Apparatus.—Same as in the last, a perforated cork to fit the test-tube, glass tubing, rubber connections.

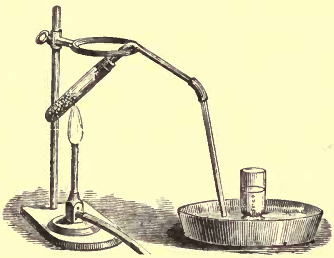


Fig. 9.

Put zinc and **HCl** as before into the test-tube, after arranging your apparatus as in the cut, except that you need no burner under the test-tube. Fill a test-tube or a bottle with water and invert in the water-pan over the end of the delivery-tube. The gas at first will be impure. Ignite and fill again. After it becomes pure it burns quietly at the mouth of the bottle. Fill one half full and then let the water out and the air in.

Hold the mouth of the bottle downwards and ignite the gas. Does it explode or burn quietly? Why?

Fill a bottle with pure hydrogen and thrust a burning splinter into the bottle. Does it burn at the surface or at the end of the splinter? Does hydrogen support combustion?

What do we mean by an explosion? Has hydrogen an odor? Color? It is not poisonous.

Experiment 38.—Hydrogen, Continued.

(*For the Teacher.*)—Generate a large quantity of hydrogen, and collect in gas-bag. Test the lightness by making soap-bubbles. Why do they rise? Do they explode when ignited? Are they pure hydrogen? Pass some of the gas through a U-tube filled with chloride of lime. Have the gas issue from a fine tube, of metal if possible, and ignite,—"Philosopher's Lamp." Why do we pass it through chloride of lime? Hold a cold bell-jar over the burning jet. What is formed upon the cold surface? Will hydrogen burn if the air is excluded? Thrust a small-mouthed bottle over the jet. Save a gas-bag of hydrogen for the next. How are balloons made to rise? Can you pour hydrogen upward from one bottle to another? Give the result of these experiments.

Experiment 39.—Synthesis of Water by Volume.

(*For the Teacher.*)—Prepare a gas-bag of oxygen, a eudiometer for exploding gases, a pan of mercury, and a battery and coil.

Fill the eudiometer with mercury and invert in the pan of mercury. Introduce equal parts by volume of each of the gases oxygen and hydrogen from the gas-bags, and allow them to mix. Pass a spark through the wires. What is the volume of the gas remaining? Reverse the tube and test the gas, and see which gas is left.

Try two volumes of **O** and one of **H**. Then two of **H** and one of **O**. Which proves to be the best proportion? Is water formed by the explosion? How does this agree with Experiment 20?

If you could weigh a litre of hydrogen accurately, you would find it about .089578 grams, and a litre of oxygen 1.429 grams. How many times heavier is oxygen than hydrogen?

According to the law of Ampère, equal volumes of all gases contain the same number of molecules under like conditions of temperature and pressure. That being true, how would the molecules of the two gases compare?

It may also be shown that the molecule of each gas contains the same number of atoms. If that is so, will the atoms compare in the same ratio? Taking **H** as the standard, as it is the lightest element, what would be the atomic weight of **O**? (See Table of Atomic Weights and Elements, page 28.) What would be the weight of a molecule of water compared with an atom of hydrogen? Name all the facts shown in the symbol $\mathbf{H_2O}$.

Since one atom of oxygen has the power of combining with two atoms of hydrogen, let us represent oxygen as having two hands with which to grasp other elements, and hydrogen but one hand. Then one atom of oxygen could grasp two atoms of hydrogen, and the hands are all occupied. In the table of elements and atomic weights you will see that these hands or combining powers are indicated thus, $\mathbf{H'}$, $\mathbf{O''}$, $\mathbf{N'''}$, $\mathbf{C''''}$. Carbon and oxygen in a satisfied union would then be $\mathbf{CO_2}$.

Write the symbol for a union of hydrogen and carbon; nitrogen and hydrogen. Call all one-handed elements Monads; two-handed, Diads; three, Triads; four, Tetrads. To aid the memory, make a table as on p. 50, putting down in its proper place each element as you learn its combining power or atomicity, or valency, as it is called. There will be only about 36, as we will not study the rare elements. Some elements, as iron, must be placed in two columns.

Monads.	Diads.	Triads.	Tetrads.
H	O	N	C

Experiment 40.—Chemistry of Air. Weight of Air.

If provided with the apparatus, weigh a litre of air.

How much does it weigh?

What apparatus is used to measure downward pressure of air? How much has this been found to be upon a square centimetre?

Experiment 41.—Oxidation of Lead.

Place a small piece of lead in a porcelain crucible, and find the weight of the crucible and its contents and of stirring-rod placed in it.

Apply heat to crucible until lead is melted, and then stir with the rod until contents takes on the form of a powder.

Weigh the three articles again, as the rod will have some particles attached to it.

Does the powder weigh more or less than the original piece of lead?

Is this an exception to the law of persistence of mass?

Where did the additional matter come from?

If you are not certain in regard to this answer, try also the following experiment:

Experiment 42.—Preventing Oxidation.

Weigh small piece of lead in crucible containing enough borax to protect lead from the air when melted. The pupil should not stir the contents of the crucible in this experiment. Does same change take place in lead as before? Why not?

Do contents increase in weight during this experiment as before? Explain.

Experiment 43.—Nitrogen of Air.

Place a piece of phosphorus, the size of a pea, on an iron sand-bath and float on water in pneumatic trough. Ignite phosphorus and cover immediately with bell-jar or large wide-mouthed bottle. Note level of water inside same as that outside at beginning of experiment. Bubbles of air escape. Explain.

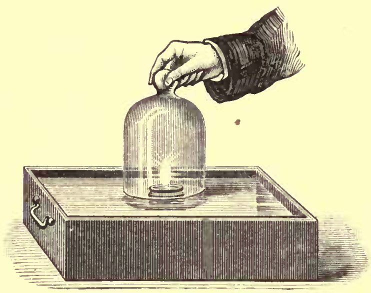


Fig. 10.—Preparing N.

Describe substance forming on inside.

Allow action to proceed and set apparatus aside until contents becomes clear. Note level of water inside.

Compare action of oxygen on phosphorus with its action on lead in previous experiments. What is color of powder in this case?

What has become of this powder? If enough phosphorus was used, it has taken up all the oxygen. If so, what proportion of original volume of air was the oxygen? How do you know this?

Experiment 44.—Nitrogen.

Apparatus.—Make a U-tube with one end drawn out to a point, leaving small opening.

Place the other end of this tube up into the jar and push the jar down into the water, forcing gas out the fine opening. Try to light this gas. Does it resemble hydrogen in this respect?

Experiment 45.—Nitrogen, Continued.

Remove tube, and cover mouth of the jar under water with a piece of glass and invert. Lower a burning taper or glowing pine splinter into gas. Does it promote combustion? Does it resemble oxygen in this respect? This gas is called *Nitrogen.*

What proportion of the original volume of air was nitrogen? (See Experiment 44.)

Experiment 46.—Per Cent. of Nitrogen.

(*For the Teacher.*)—Mix 100 volumes of air and 50 volumes hydrogen in a eudiometer and explode as in synthesis of water. Water formed is condensed and shrinkage in volume noted is 63—*i.e.*, volume at beginning, 150; after exploding, 87: thus, $150 - 87 =$ loss of 63. Of this, of course $\frac{1}{3}$ is oxygen and $\frac{2}{3}$ hydrogen. Hence, $\frac{63}{3} = 21$ parts of oxygen in 100 parts of air.

Experiment 47.—Analysis of Air.

By Prof. Le Roy C. Cooley, Ph.D., in "A Guide to Elementary Chemistry for Beginners."

We set out now to find how many cubic centimetres of nitrogen and how many of oxygen and carbon dioxide there are in 100 c. c. of air.

To do this we will imprison a vesselful of air, and then run into it a liquid which will absorb both the oxygen and the carbon dioxide completely, and leave the nitrogen. We can then measure the nitrogen which is left, and we can find out how much there was of the other two by measuring the liquid which had gone into the tube to take their place.

Our Apparatus.—I take a test-tube, *t* (Fig. 11), to hold the air. A six-inch tube, $\frac{5}{8}$ inch in diameter, will do; an eight-inch tube of the same diameter is better. The rubber stopper, *c*, is so large that its small end will enter the tube only about a half-inch. It has two holes; to close one I have a solid rod of glass, *s*; for the other, a glass tube reaching just a very little below the cork, as shown. A piece of thin rubber tubing, *h*, is cut about six inches long. There is a pinch-cock, *p*, by which its walls may be pinched so as to close it completely. *F* is a small glass funnel.

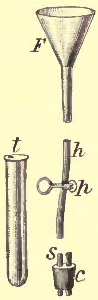


Fig. 11.

The lower end of *h* I stretch over the tube in the cork *c*, and its upper end I fix over the stem of *F*, and then I place the funnel in the clamp of the support, as shown in Fig. 12, and remove the rod *s*.

The Liquid.—To absorb the oxygen and carbon dioxide gases I use a mixture of pyrogallic acid and potassium hydrate.

I take a small teaspoonful of the solid acid and pour on it 10 c. c. of water; it will soon dissolve. To this I then add 5 c. c. of strong solution of potassium hydrate, and *at once* pour it into the funnel. Next, I hold the dish below the cork and open the pinch-cock *p* a moment, to let the liquid run down and fill the tubes completely. I carefully take off the drop, which hangs at the lower end of the tube below the cork, with a piece of filter-paper.

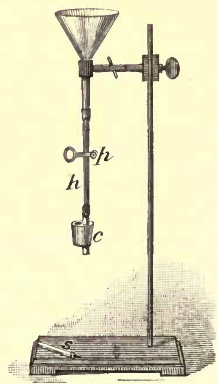


Fig. 12.

I press the tube *t* up over the cork until the joint is air-tight, as seen in Fig. 13, and after a minute I put the rod *s* into the open hole of the cork. I have now imprisoned a tubeful of air; none can get out, and no more can get in.

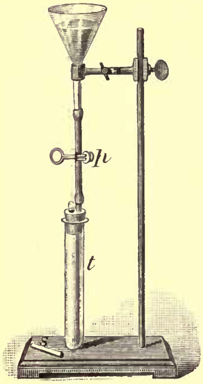


Fig. 13.

I left the hole in the cork open, because if it were not open the pressure of the cork would crowd the air below, and there would be too much in the tube; and then, too, handling the tube warmed it, and the volume of air changes with heat. With the hole open, the air in the tube soon comes to be just as warm and just as much pressed as the air outside. Whenever a gas of any kind is to be measured, its *temperature* and *pressure* must be the *same as those of the air outside.*

The Absorption.—I now press the pinch-cock *p*; a little stream of the liquid falls into *t* at once, and then drops follow, or, if the tube be slightly inclined, a slender stream will flow down its side. It will continue to enter as long as there is any oxygen or carbon dioxide for it to absorb, and then stop. The gas which is left in the tube is nitrogen.

But this gas is crowded down by the pressure of the liquid in the rubber tube and funnel above, and so I take hold of the cork *c*, and the rim of *t*, not to warm the gas with my hand, and lift the tube bottom up, as shown at *T* in Fig. 14, making the level of the liquid the same in the tube and in the funnel. I then open the pinch-cock. Some of the liquid will run out of *T*. When the liquid stands at the same level in the tube and in the funnel, I close the cock and bring the tube down again.

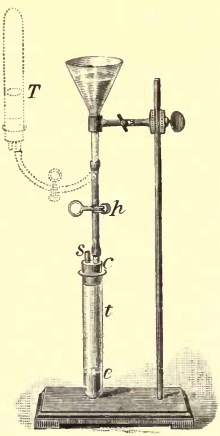


Fig. 14.

The almost black liquid in *t* has now taken out all the oxygen and carbon dioxide from the tubeful of air, and left all the nitrogen.

The Measuring.—I must measure the liquid in the tube to find how much oxygen was taken out—and carbon dioxide also; but the volume of the carbon dioxide, in so small a quantity of air as we use, is so little that we cannot measure it with our apparatus, and therefore leave it out of account in this experiment—and the space above it to find how much nitrogen was left. To do this I slip two small rubber rings upon the tube, and make the upper edge of one mark the place of

the lower end of the cork, and of the other, the top of the liquid. These rings must not afterward be disturbed.

I may now remove the cork, empty the tube, rinse it with water, and then let the last drop of water drain away. Finally, I use my graduated cylinder to find out exactly—

How many c. c. of water will fill the tube to the first ring?

How many c. c. from the first to the second ring?

The Calculations.—From these two numbers we can find what part of the air is nitrogen and what part is oxygen, for they help us to answer the following questions, in their order, as shown by an example below:

How many c. c. of air were in the tube at first?

How many c. c. of nitrogen did this air yield?

How many c. c. of oxygen did the same air yield?

Then what fractional part of the air is nitrogen?

What fractional part of the air is oxygen?

And how many c. c. nitrogen in 100 c. c. of air?

How many c. c. of oxygen in 100 c. c. of air?

An Example.—In an actual experiment it was found to take of

Water to fill the tube to the first ring, ...	6.0 c. c.
Water to fill the tube from the first to second ring, . .	23.5 c. c.
Hence the number of c. c. of air taken, ...	29.5 c. c.
And the number of c. c. of nitrogen found, . . .	23.5 c. c.
And the number of c. c. of oxygen found, ...	6.0 c. c.

Now this would show plainly that $\frac{235}{295}$ of the air is nitrogen and $\frac{60}{295}$ of it is oxygen. Then in 100 c. c. of air there would be

Nitrogen, ...	79.66 c. c.
Oxygen, ...	20.34 c. c.

Experiment 48.—Moisture in Air.

Does air contain any other gas besides **N** and **O**? Breathe into a cold, clean flask. What is found on the inside? Where does this come from? Is there some of this substance in the air all of the time? To answer this question, think of the collecting of water on sides of ice-pitcher in summer. Explain this. What is dew? What is frost?

Experiment 49.—Carbon Dioxide in Air.

Put some lime-water into the flask, breathe into it again, shake up contents of flask and note result, or use beaker of lime-water and blow into it through glass tube. The gas which produces this result is called *Carbon Dioxide.* Where does it come from in this experiment? How many substances have you now seen to be in the air? Is their proportion always the same? Is air a chemical compound or a mechanical mixture?

Experiment 50.—Preparation of Oxygen.

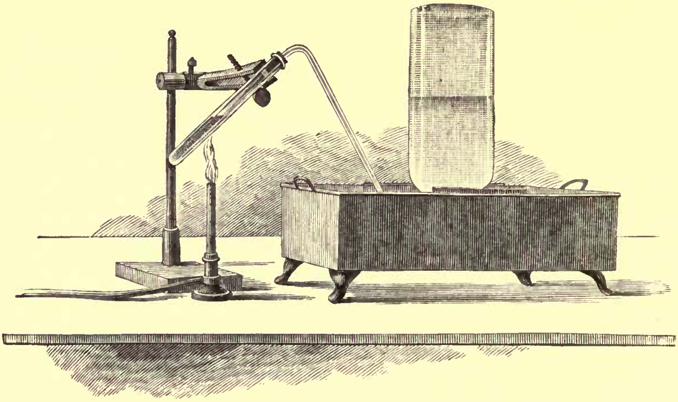


Fig. 15.—Preparation of Oxygen.

Make a quantity of oxygen as directed in Eliot and Storer's "Elementary Manual of Chemistry," Experiment 4, page 9.

Experiment 51.—Characteristics of O.

(*For the Teacher.*)—Let the teacher perform the experiments given in the text-book, or others, if he prefers, to illustrate the properties of oxygen. Pupils should note each step in each experiment and write out carefully what is shown by it.

Experiment 52.—Effect of O on Blood.

N.B.—The teacher should obtain, from the nearest slaughter-house, a quantity of freshly-drawn blood, which has been whipped to take out the fibrine so that it will not coagulate. This may be kept for several days, if necessary, in a cool place, in a well-corked bottle.

Take small quantity of the blood in a test-tube. Notice its dark-red color. This shows you the color of so-called venous blood.

Now shake up the test-tube, placing the thumb over the top.

After mixing air in tube with the blood, remove the thumb and admit more fresh air, and shake as before.

Repeat this several times, and note change in color of the blood.

Where in our bodies does this change of color take place? What part of the blood gives it its red color?

Explain the change which takes place in the blood in the lungs.

What change takes place in the air which goes to the lungs?

Experiment 53.—Oxygen from Green Plants.

(*For the Teacher.*)—A tall, wide-mouthed bottle, water-pan, water-plants.

Fill a tall, wide-mouthed bottle half full of water-plants. The light-green fibrous alga *spirogyra* is the best. Fill the bottle with water and invert in the water-pan, and support it so there can be circulation of water from the bottle to the pan. Place the apparatus in the sun-light and allow it to stand until a sufficient quantity of gas is generated. Cover the mouth of the bottle and invert. Test for oxygen with a burning splinter. Where did the bubbles of gas come from? What had the light to do with it? What effect have water-plants upon stagnant pools? Do plants in the air do the same work? What is the gas in the water that chlorophyll decomposes? How does it come there? How are plants and animals related in respect to the use of oxygen? Could the number of plants and animals in an aquarium be so balanced that each would furnish the gaseous food required by the other?

Experiment 54.—Nitrous Oxide.

Compounds of **N** *and* **O.**—Place a small quantity of ammonium nitrate ($\mathbf{NH_4NO_3}$) in a test-tube and heat.

Hold piece of cold glass near mouth of tube, and note what collects upon it.

What is the first change which takes place in the contents of the tube? What is the next? Which is chemical? Is there any residue?

Complete the equation, $\mathbf{NH_4NO_3} = 2\mathbf{H_2O} +$.

Experiment 55.—Nitrous Oxide, Continued.

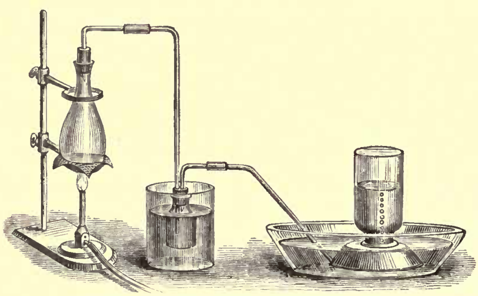


Fig. 16.—Preparing Nitrous Oxide.

Arrange apparatus as shown in Fig. 16.

Test gas collected in bottle with glowing splinter, as in experiment with oxygen. What is the result?

Experiment 56. — Nitric Oxide.

N.B.—To be performed under the hood or where there is a good draught.

In flask used for generating hydrogen place a few grams of sheet-copper cut in small pieces, and connect delivery-tube to wide-mouthed bottle in pneumatic trough as in preparing hydrogen. Pour in enough 50 per cent. nitric acid to cover the copper and to come above the opening of the thistle-tube. (*Do not inhale the fumes.*)

Collect the gas given off, by displacement of water, as in the case of hydrogen. What is in the generating flask before the beginning of the experiment? What is the color of the gas given off at first? What is the color after a quantity has been collected over water?

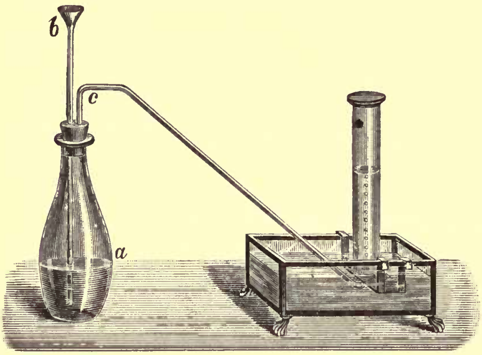


Fig. 17.—Preparing **NO**.

Note. — As soon as a small quantity of the gas has been collected, prevent further formation of it, by disconnecting the generating flask, and pouring off liquid which it contains, into an evaporating dish and setting it aside. Rinse copper quickly in plenty of water and put away to be used again.

Experiment 57.—Nitric Oxide, Continued.

Allow a small amount of the gas which you have collected to escape into the air. What is the color when it comes in contact with the air? What element in air is most active and most likely to be the cause of this change? Does the gas after it has passed into the air resemble that in the generating flask when action first commenced? How do you account for this? Complete the equation,

$$3Cu + 8HNO_3 = 3Cu(NO_3) + H_2O + NO.$$

This represents the chemical action in making nitric oxide (**NO**).

$$NO + O = NO_2 \text{ (nitrogen peroxide).}$$

How many oxides of nitrogen are described in your text-book? Make a table of them, showing relation of **N** to **O**, by weight in each.

What fundamental law of chemical action may be derived from a consideration of these compounds?

What experiments have we made use of in reaching the same law?

Evaporate a small portion of the liquid. Set aside in evaporating dish. What remains? Name of the substance? Its formula?

Experiment 58.—Compound of N and H.

Mix five grams of ammonium chloride and same weight of cold, freshly slaked lime, $Ca(OH)_2$. Arrange flask as shown in cut. Note odor of gas given off. Warm gently for half a minute and try to collect gas by displacement of water as in case of **NO** and **H**. Do you succeed in thus collecting the gas? Test the water with litmus. Is the liquid acid or alkaline? What is the liquid?

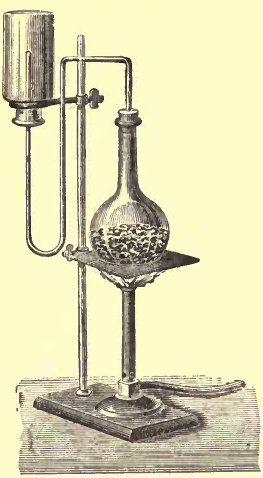


Fig. 18.—PREPARING NH_3.

Experiment 59.—Ammonia from Ammonium Chloride.

Note odor of ammonium chloride. Mix small quantity with lime as in preceding experiment. What is the odor? This substance is called ammonia (NH_3). The complete equation is,

$$2NH_4Cl + Ca(OH)_2 = 2NH_3 + CaCl_2 + 2H_2O,$$

also $NH_3 + H_2O = NH_4HO$.

Put some solution of ammonium chloride in a test-tube and add a few drops of **KOH**. Result? Write the equation.

Experiment 60.—Chlorine.

Materials and Apparatus.—Test-tube, pulverized metallic antimony, black oxide of manganese, hydrochloric acid, piece of calico.

Put a gram of oxide of manganese into a test-tube and cover with **HCl**. Heat gently, being careful not to breathe any gas that may be given off. What is the color of the gas? Hold a piece of wet calico over the mouth of the test-tube so the escaping chlorine will have to pass through it. How does it affect the color? Try dry calico. Write upon a piece of paper with carmine ink and hold it over the test-tube. Is it bleached? Try common print moistened. Why is it not bleached? Incline the tube and see if the gas falls or rises as it escapes. Drop some powdered antimony into the tube as the gas is rising freely. What occurs?

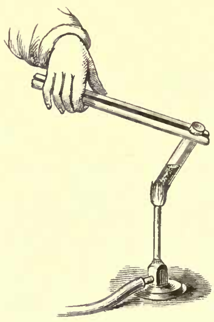


Fig. 19.

Complete the equation, $MnO_2 + 4HCl = MnCl_2 + 2H_2O +$?

Again, $Sb + 3Cl =$? What is the cause of the fire seen in the test-tube? Of what is the white smoke composed?

Experiment 61.—Chlorine, Continued.

(*For the Teacher.*) — Let the teacher prepare chlorine in a large quantity, and apply the following tests. Fill several bottles with the gas and pass the gas through water in Wolfe bottles to prepare chlorine water.

Wet a strip of paper in warm water and thrust it down into a bottle of gas. Does it show chemical action?

Lower a lighted taper into a bottle of gas. Will it burn?

Light a ribbon of manganese and lower it into a bottle of gas. Will it burn?

Set a bottle of chlorine water inverted in a pan of chlorine water in the sun-light. Examine after twenty-four hours. Has gas collected?

What is it? Where did it come from?

Insert a cork, with a short piece of tubing through it, into a bottle full of chlorine. Invert this over a beaker of water. After half an hour or so examine. Why has water partly filled the bottle?

Put 10 grams of chloride of lime into a litre of water; add a few drops of **HCl**.

Immerse calico, etc., to see if it bleaches readily. Will chlorine water remove fruit stains?

Experiment 62.—Chlorine Water.

Drop a few crystals of potassium chlorate ($KClO_3$) into a test-tube and add hydrochloric acid. Warm gently, and when the yellowish gas begins to appear, add 4 or 5 c. c. of water.

As fast as chlorine is set free it will be absorbed by the water. The reaction is

$$4HCl + 2KClO_3 = 2KCl + 2H_2O + Cl_2O_4 + 2Cl.$$

What effect has this water upon vegetable coloring-matter, as litmus, carmine, etc.?

In nearly all cases where bleaching is done by chlorine the reaction is

$$2Cl + H_2O = 2HCl + O.$$

The oxygen set free burns up the coloring-matter, so that in reality oxygen does the bleaching.

In what form is chlorine sold by the druggist for bleaching?

Write a full account of the preparation of chlorine, giving its chief characteristics, and the uses you see that can be made of it.

Experiment 63.—Bromine.

Materials and Apparatus.—Test-tube, bromide of potassium, manganese dioxide, sulphuric acid, chlorine water.

Mix half a gram of potassium bromide and a gram of manganese dioxide in a test-tube, and cover with dilute sulphuric acid. Heat it. Do not breathe the gas. How does bromine differ from chlorine? Will it bleach calico or litmus paper? Compare the strength of chlorine and bromine by the following: Make a solution of a few crystals of **KBr** in two or three c. c. of water; add a drop or two of chlorine water. Is **Br** set free? Explain the reaction, $\mathbf{KBr + Cl = KCl +}$ (free) **Br.**

Has chlorine more intense chemical action than bromine?

Experiment 64.—Iodine.

Materials and Apparatus.—Test-tube, crystals of iodine, potassium iodide, alcohol, chlorine water, starch or flour.

Examine a crystal of iodine. Does it stain the fingers? Dissolve in water. Make a thin paste of starch or flour and add a drop of the solution. What occurs? Dissolve half a gram of potassium iodide (**KI**) in two or three c. c. of water. Add to this a c. c. of starch paste. Does it stain as the element iodine? Why not? Now add some chlorine water. Does it now? Complete the equation, $\mathbf{KI + Cl = KCl +}$? Does this show that chlorine is stronger than iodine? Does it bleach? How can you detect the presence of starch? How can you distinguish between these three elements? How are they related? Dissolve a crystal of iodine in alcohol. This is a *tincture.* Allow it to evaporate. Does iodine recrystallize?

Experiment 65.—Hydrochloric Acid, HCl.

Materials and Apparatus.—Test-tube, litmus paper, common salt, silver nitrate, ammonium hydrate, perforated stopper and tubing, beaker.

Put 2 c. c. of common salt into a test-tube and add 2 c. c. of concentrated sulphuric acid. Is a gas given off? Test with a strip of moistened blue litmus. Is it an acid? Complete the reaction, $\mathbf{2NaCl + H_2SO_4 = Na_2SO_4 +}$? Hold an open ammonium-hydrate bottle near the mouth of the test-tube. What occurs? Now hold the ammonium-hydrate and hydrochloric-acid bottles unstoppered near each other. Does the same occur?

$\mathbf{HCl + NH_3 = NH_4Cl.}$ What is the difference between the reagent **HCl** and the gas you have generated? Insert the stopper and pass the gas into a beaker or bottle of water for some time. Does this water acquire acid properties? Is sodium sulphate left in the test-tube after the bubbling ceases? What series of salts is formed from **HCl**?

Experiment 66.—Test for HCl.

Take two or three drops of $AgNO_3$ in a test-tube, and add a few drops of **HCl**, and we have a white precipitate of silver chloride, thus, $AgNO_3 + HCl = AgCl + HNO_3$.

Now add ammonium hydrate and it will all dissolve.

Now take any salt of hydrochloric acid, as NH_4Cl, $CaCl_2$, or $BaCl_2$, and add $AgNO_3$.

Do you get the same white precipitate of silver chloride?

Try its solubility in NH_4HO. By this means you can tell any *chloride.*

Experiment 67.—Nitric Acid, HNO_3.

Let the teacher prepare nitric acid as in Eliot and Storer's "Elementary Chemistry," page 39, Experiment 22.

Pupils try the following characteristic test for nitric acid and nitrates:

Make a solution of a c. c of ferrous sulphate ($FeSO_4$) in 10 c. c. of water. Add to this 2 c. c. of H_2SO_4 and allow to cool.

What caused the heat?

Now incline the test-tube so that you can pour in two or three c. c. of dilute nitric acid without its mixing with the heavy solution below.

Where the two liquids meet, a brown ring will be formed.

Now do the same again, except that instead of using nitric acid use a solution of any salt of nitric acid, as KNO_3 or $NaNO_3$.

Does it act in the same way?

By this test you may tell the presence of nitric acid or a nitrate. Describe the manufacture of HNO_3.

Experiment 68.—Sulphuric Acid, H_2SO_4.

The manufacture of H_2SO_4 is too difficult to perform well. Characteristic test:

Add $BaCl_2$ to sulphuric acid, and a heavy white precipitate of $BaSO_4$ will be formed.

Add *aqua regia*—a mixture of concentrated **HCl** and HNO_3—and it will not dissolve.

Try any sulphate, as $FeSO_4$ or $CuSO_4$, and see if they act the same as H_2SO_4.

Complete these equations:

$$BaCl_2 + H_2SO_4 = BaSO_4 + ? \quad BaCl_2 + CuSO_4 = ? + ?$$

Experiment 69.—Hydrosulphuric Acid, H_2S.

Place a small fragment of ferrous sulphide (**FeS**) in a test-tube and add 4 or 5 c. c. of water. Now add a c. c. of sulphuric acid.

As chemical action begins notice the odor of rotten eggs. This acid is a gas, but may be passed into water and the saturated water used as with **HCl**.

Complete the equation, $FeS + H_2SO_4 = H_2S + ?$

Moisten a piece of filter-paper with lead acetate and hold it over the mouth of the tube.

Is it blackened?

Do not allow it to generate gas very long, as it is poisonous. Rinse out the tube, saving the **FeS** for future use.

Characteristic Test.—Now take a solution of a salt of this acid, as ammonium sulphide, $(NH_4)_2S$, in a test-tube. Add **HCl**.

Is H_2S given off?

Test with lead acetate as before.

Not all sulphides will be shown by this test; but if first heated on charcoal with Na_2CO_3 and then heated with **HCl**, they yield H_2S.

Can you generate H_2S with **HCl** and **FeS**? Try it.

For generating H_2S on a larger scale, see Eliot and Storer's "Elementary Chemistry," page 76.

Experiment 70.—Carbonic Acid, H_2CO_3.

This acid does not exist uncombined, as it splits up into two parts, thus:

$$H_2CO_3 = H_2O + CO_2.$$

Its salts are abundant, and are detected by the following:

Take any carbonate in a test-tube and add **HCl**, and rapid effervescence will occur.

While the test-tube is filling with gas, fill another test-tube $\frac{1}{4}$ full of lime-water, CaH_2O_2.

Incline the test-tubes as if you would pour a heavy gas from the first tube into the lime-water. After a moment cover the mouth of the latter tube with the thumb and shake thoroughly.

Does the water become clouded? What gas must it be?

Characteristic test of carbonates, except that oxalates may yield CO_2. The oxalates, however, when heated on a platinum-foil leave a black mass and carbonates do not.

NOTE.—Let the teacher give several salts of the acids given here to be distinguished by these tests.

Experiment 71.—Sulphuric Dioxide.

Materials and Apparatus.—Deflagration-spoon, sulphur, large bottle, a red rose, nitric acid, some cotton, sulphuric acid.

Light a small lump of sulphur in the deflagration-spoon and lower it into a large wide-mouthed bottle covered by a card. As sulphur burns what gas is formed? Compare the valency of the two elements. What does the odor of the gas resemble? Will it burn? Try a lighted splinter. Hold a red rose or any red flower in the mouth of the bottle. What effect? Fill a beaker with water and pour a little sulphuric acid into it.

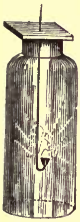


Fig. 20.

[NOTE.—Never pour water into sulphuric acid, but pour the acid into water. Why?] Immerse the rose in the dilute acid. Dry and warm it. Is the color restored? Try to bleach wet calico, also blue litmus paper. What effect? Does sulphur dioxide bleach as chlorine does, by combining with the hydrogen of water, setting free the oxygen, the latter burning up the coloring-matter? Straw hats are bleached by sulphur. Why do they become yellow when wet? Why do flannels become yellow? By means of a wire or glass rod lower a piece of cotton or sponge wet with nitric acid into the bottle of gas. The red fumes are the result of reduction or the taking away of oxygen from the acid, thus: $\mathbf{2HNO_3 + SO_2 = H_2SO_4 + 2NO_2}$. Give the name of each as indicated.

Does this suggest what might be used in making sulphuric acid?

Experiment 72.—Phosphorus.

Materials and Apparatus.—Phosphorus, carbon bisulphide, beaker, filter-paper.

Put a lump of phosphorus as large as a grain of wheat into a beaker and pour upon it immediately a c. c. of carbon bisulphide. Does it dissolve? Pour a few drops of the solution upon a filter-paper and allow it to evaporate. The paper may be held upon the palm of the hand if care is taken not to hold it too near the person, so that the bursting into flame will not set fire to the clothing. Why does it burn spontaneously? Why does not the phosphorus on the head of a match burn spontaneously? What use have you seen made of phosphorus in preparing other elements? Why is phosphorus kept under water? Why does it not burn up the oxygen contained in the water?

NOTE.—Teacher perform Experiment 59, Eliot and Storer's "Elementary Chemistry," page 96.

Experiment 73.—Arsenic.

Materials and Apparatus.—A hard glass tube, some charcoal, and white arsenic.

Seal up one end of a piece of glass tubing four inches long. Drop into this a little white arsenic, then cover with some small lumps of charcoal. Heat the charcoal red-hot first, then heat the arsenic below it. What is the effect?

What power has carbon when hot with reference to oxygen in other compounds at hand?

What is the metallic substance above the charcoal?

Experiment 74.—Test for Arsenic.

Materials and Apparatus.—Test-tube, zinc, sulphuric acid, sodium hydrate, mercuric chloride, filter-paper.

Put some grains of zinc into a test-tube and cover with dilute sulphuric acid.

Add a solution of white arsenic, made by boiling the powder in sodium hydrate and water.

As hydrogen escapes from the acid by the action of zinc, it will combine with the arsenic and escape as a gas, $\mathbf{AsH_3}$, arseniureted hydrogen.

Cover the mouth of the test-tube with a filter-paper moistened with mercuric chloride, and in a few moments the paper will be stained yellow to orange red. (*Characteristic test.*)

Experiment 75.—Carbon Dioxide.

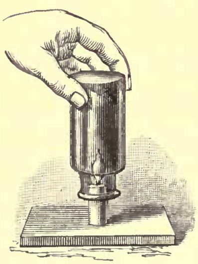


Fig. 21.

Place a short piece of candle on your desk; light, and hold a dry, clean, wide-mouthed bottle down over the flame.

What do you see formed on the inner wall of the bottle?

Remove bottle, pour in a small quantity of lime-water, and shake.

What color is given to the lime-water?

In what experiment has this change been produced before?

What substance is formed?

Write the equation.

What two compounds are seen to be the products of combustion? Review Experiments 49 and 50.

Experiment 76.—Preparation of CO_2 for Use.

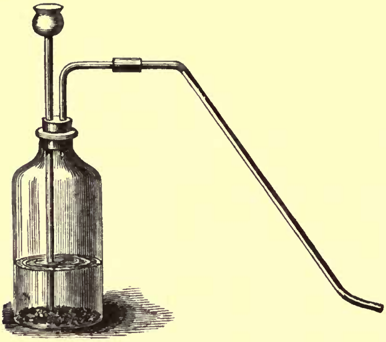


Fig. 22.

Place several small lumps of marble or lime-stone (calcium carbonate) in generating flask used for making **H**. Cover with water, and pour in strong **HCl** until gas is given off abundantly. Collect in jars over water, as in the case of hydrogen. Collect some in wide-mouthed bottle by displacement of air, having bottle on your table with mouth up, and delivery-tube running to the bottom.

Complete the equation, $\mathbf{CaCO_3 + 2HCl = ?}$ Try the effect of some of this gas upon lime-water. Name of the gas? Stand short bit of lighted candle in bottom of a tumbler and pour gas upon it as you would water. What is the effect upon the flame? What does this last action show as to weight of this gas? Find the weight of this gas, if you have not done so already, by balancing a large beaker on the scale-pan and filling with $\mathbf{CO_2}$ by displacement of air.

How would you pour **H** from one vessel to another containing air? In what position would you hold the bottles?

Fig. 23.—Weighing Carbon Dioxide.

Experiment 77.—Absorption of Carbon Dioxide.

Fill a bottle half full of $\mathbf{CO_2}$ over water in a pneumatic trough. Cork the bottle under water and shake. Lower mouth of bottle into water again, remove stopper, and note the rise of water on the inside. To what is this due?

What is "soda-water"? What other gas which you have studied was dissolved in water?

Experiment 78.—Saturation with Carbon Dioxide.

Conduct some CO_2 from delivery-tube into beaker of lime-water. Note whitening as before. What substance is formed? Is it soluble in water?

Continue to pass the gas until white color seems to disappear. Can you explain? What does water contain after the gas has passed through it? What is the effect of water in this condition upon limestone?

Heat the contents of the beaker and see if the white substance reappears. When some water is boiled in a tea-kettle, limestone is deposited upon the bottom and sides. Explain. What gas does such water contain in solution before boiling? In caves and limestone regions the water, in coming to the surface, deposits limestone in the form of icicles, some hanging down from above and others rising from the floor. Can you give the explanation?

How does dough rise? (See Experiments 16 and 90.)

Experiment 79.—Candle Flame.

Examine the flame of a candle. Have it as large as possible, and protected from draughts of air. Note (1) the blue cup-shaped portion at the base, (2) the dark, non-luminous portion inside surrounded by the (3) luminous cone, and (4) outside of that by a thin layer or region of complete combustion. Make a sketch of the flame in your note-book, showing all its parts.

NOTE.—The spectrum given by a solid or liquid in a state of incandescence is a continuous band of color. That given by a gas in a burning state consists of bright lines here and there across the place occupied by the continuous spectrum when present.

Examine the candle flame with a spectroscope. Does it show the light-giving portion of the flame to consist of a solid or liquid, on the one hand, or of a gas?

Hold a cold iron spoon down upon the flame until it almost touches the wick. What collects upon the spoon? Next hold this blackened surface for some time in the non-luminous flame of the Bunsen burner, and describe what takes place. What is your conclusion as to the condition of matter in the luminous portion of a flame? At which portion of the flame does the air have freest access to the combustible material? Where is the region of complete combustion? Is this luminous or non-luminous? Press down upon the flame a piece of white paper or cardboard. Hold it there for an instant and remove quickly to avoid taking fire. What does this show in regard to the interior of the flame? Where was the paper charred? Sketch cross-section in your notes.

Experiment 80.—The Flame.

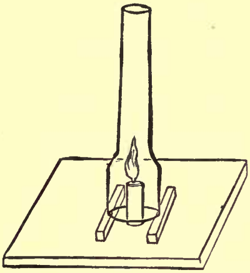


Fig. 24.

(*a*) Stand a lighted candle on smooth table, and place a tall, narrow lamp-chimney over it, so that no air can enter from below.

Does the candle continue to burn?

(*b*) Now raise edges of chimney up on thin strips of wood or metal so that air may enter. What is shown by this experiment?

Now, with candle as in this experiment (*b*), cover chimney with strips of metal or evaporating dish. Explain what happens.

Experiment 81.—Kindling Temperature.

Lower spoon in a candle flame almost to the wick, and collect carbon upon it. Why is carbon deposited, and why does it not continue to burn up? What is "kindling temperature"?

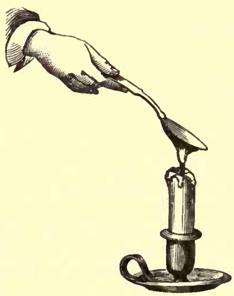


Fig. 25.

Fold a piece of commercial note-paper as you do your filter-paper. Pour a little water into it, and place in one of the rings of your ring-stand, and boil water with Bunsen burner.

What is temperature of boiling water? What must be the temperature of the paper next to it? Is this temperature as high as the kindling temperature of the paper? Explain the purpose of using a match, pieces of paper, and wood, in this order, in building a coal fire.

Experiment 82.—The Bunsen Burner.

When the holes are closed at the bottom of the lamp, light the gas and examine the flame. Compare with candle flame. Explain the change in character of the flame produced, with the orifices at the bottom of the lamp open. Find the hottest part of this flame by means of a platinum wire. Lower a piece of wire-gauze down upon the flame, and notice structure of the flame in cross-section by seeing where the wire-gauze is made to glow.

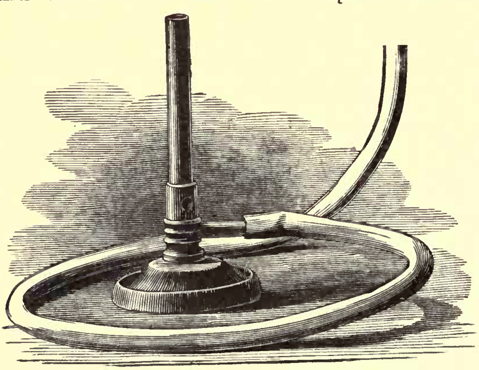


Fig. 26.—Bunsen Burner.

Experiment 83.—Principle of Sir Humphry Davy's Safety-lamp.

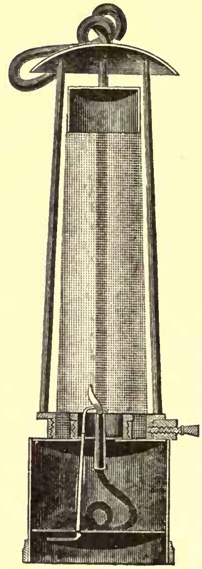


Fig. 28.—Davy's Safety-lamp.

Arrange wire-gauze on ring-stand, as shown in Fig. 27. Turn on the gas and light it above the gauze. Is there gas below the gauze? Prove it.

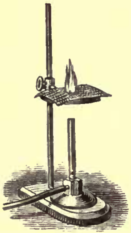


Fig. 27.

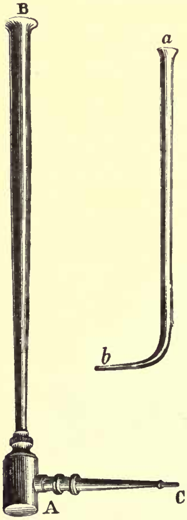
Extinguish flame and light the gas below the gauze. Does the flame pass through it? Explain.

Look up the subject of "Construction and Use of Davy's Safety-lamp," page 134 of Eliot and Storer's "Elementary Chemistry," and elsewhere.

Experiment 84.—Action of Hot Carbon upon Oxygen in Combination.

Mix thoroughly .25 gram powdered charcoal and 2.5 grams copper oxide, and place in a hard glass test-tube provided with a delivery-tube.

Heat persistently over Bunsen flame, and conduct gas into another test-tube containing lime-water. What is the gas? After gas ceases to be given off, examine contents of hard glass tube. Is it copper? By weighing tube before and after heating, find weight of copper in 2.5 grams of copper oxide.



Experiment 85.—The Blow-pipe.

Note.—Read carefully Eliot and Storer's "Elementary Manual of Chemistry," pages 129 and 130.

Place a small quantity **PbO** in a hollow made in piece of charcoal. Heat in the "reducing flame" for some time, until minute drops of metal are seen in place of the powder. What is the metal? Try to cut it with a knife. Write equation for above reaction.

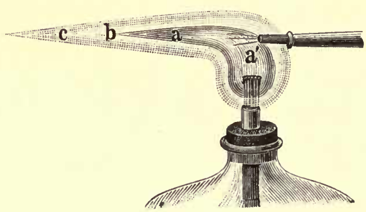


Fig. 29.—Common Blow-pipes.

Note.—Experiments 84 and 85 illustrate in general the method of securing metals from their ores, as metallic iron from iron oxide.

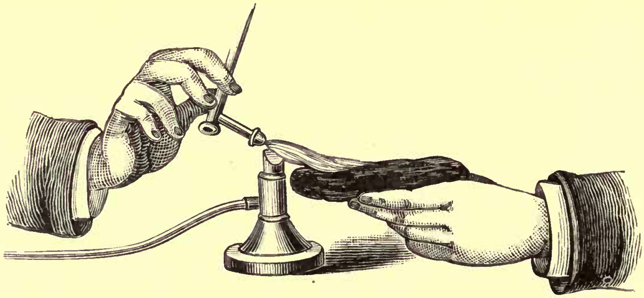


Fig. 30.—USE OF BLOW-PIPE.

Experiment 86.—The Oxidizing Flame.

Place a small bit of lead on charcoal hollowed out as in preceding experiment, and heat in the oxidizing flame. What is the name of the coating found about the lead on the charcoal? What substance does it resemble? Write the equation. Has the same substance been produced before? In what experiment?

Experiment 87.—The Borax Bead.

Follow the directions for making a borax bead given in Eliot and Storer's "Elementary Chemistry," pages 217 and 218. Touch the bead to a small speck of black oxide of manganese, and heat in the oxidizing flame. Pulverize this bead between heavy pieces of iron. Be careful not to bend the platinum wire. Clean the wire and make a new bead. Dip this in cobaltous nitrate solution, or in the salt itself, and heat as before. Note color when hot and when cold.

Borax-bead tests for other metals will be referred to in Experiments on Metals, further along in this book.

NOTE.—Borax dissolves the oxides of various metals; hence its use in preparing a clean surface in soldering.

Experiment 88.—Marsh Gas.

(*For the Teacher.*)—Perform the experiment as directed on page 138 Eliot and Storer's "Elementary Chemistry."

Why does it explode when mixed with air and ignited? In what proportions do they explode best? What must be formed when CH_4 burns? How can you prove that water is one of the products? CO_2? What proof that carbon is left when chlorine and CH_4 explode? Why do miners call this gas *fire-damp?* Why does Davy's safety-lamp prevent an explosion? Under what conditions will an explosion occur? This gas forms a large part of illuminating gas. Try to collect the gas from a pond. Describe the whole experiment.

Experiment 89.—Illuminating Gas.

Materials and Apparatus.—Two clay pipes, plaster of Paris, soft coal, sawdust.

Fill a clay pipe half full of small pieces of soft coal. Stop the mouth of the bowl with a paste made of plaster of Paris and water. As soon as the plaster hardens support the pipe so that it can be strongly heated. At first, steam will issue from the stem of the pipe, and finally, gas. Ignite from time to time until the gas becomes pure enough to burn of itself.

After all the gas is driven off let it cool, then cut out the plaster and examine the contents of the pipe. What is the black mass? Do the same with some sawdust or fine chips. Do they yield a gas that will burn? What is left after the burning or heating is finished? How does coke differ from charcoal? Why did not the wood burn to ashes? How is charcoal prepared for commerce? Some of the impurities formed when coal is thus heated to make illuminating gas are $\mathbf{CO_2}$ and $\mathbf{NH_3}$. How could these two impurities be removed? Does $\mathbf{CH_4}$ from the pipe-stem burn with a clear or smoky flame? Is carbon wasting? How can this be remedied? Visit some gas-works and see how gas is made on a large scale.

Experiment 90.—Alcoholic Fermentation.

Dissolve 30 grams of grape-sugar ($\mathbf{C_6H_{12}O_6}$) or cane-sugar ($\mathbf{C_{12}H_{22}O_{11}}$) in 400 c. c. of water. Put the solution into a flask provided with a stopper and delivery-tube to conduct the gas generated into a bottle of lime-water ($\mathbf{CaH_2O_2}$).

Add to the solution of sugar a fourth of a cake of compressed yeast or a tablespoonful of baker's yeast. Insert the stopper and place the apparatus in a warm place for from twenty-four to forty-eight hours. Does the lime-water indicate that $\mathbf{CO_2}$ has been generated? Taste the solution. Is it sweet? The chemical change may be indicated by the equation, $\mathbf{C_6H_{12}O_6 = 2C_2H_6O + 2CO_2}$. That is, by the action of yeast, grape-sugar splits up into two molecules each of alcohol and carbon dioxide.

Or if cane-sugar was used it changes first to grape-sugar, thus: $\mathbf{C_{12}H_{22}O_{11} + H_2O = 2C_6H_{12}O_6}$.

From what source is alcohol obtained? Since the boiling point of alcohol is 20° lower than that of water, can you suggest a method of separating alcohol from water?

Save the dilute alcohol you have made for the next experiment.

Experiment 91.—Vinegar.

Allow the dilute alcohol made in the last experiment to stand several days exposed to the air. Finally, it will acquire the characteristic taste of vinegar. The reaction is: $C_2H_6O+O_2=HC_2H_3O_2+H_2O$. From what source is vinegar (acetic acid) obtained? Sometimes water is made sour by a small quantity of sulphuric acid. How could you detect it? (See Sulphuric Acid, page 72.) What is yeast? Is the change from alcohol to vinegar due to yeast? How can you prove it?

Experiment 92.—Fractional Distillation.

(*For the Teacher.*)—Collect part of the dilute alcohol from the class or prepare some in a similar way. Arrange apparatus as shown in Eliot and Storer's "Elementary Chemistry," page 146. Distill several times to get sufficient strength of alcohol to burn. Put a few drops of alcohol upon the hand. Why is it cold? Drop some camphor-gum into a few c. c. of alcohol. Does it dissolve? Add water and it will be precipitated. Pour strong alcohol upon white of egg. What effect? Put some green leaves into a bottle and cover with alcohol. Allow to stand over night. Is the green coloring matter (chlorophyl) extracted? Mix together two volumes of concentrated alcohol and two of water. Does it make four volumes of the mixture? How do you account for it? How does alcohol preserve museum specimens?

Experiment 93.—Test for Sugars.

Dissolve rock-candy or granulated sugar in water. Take 5 c. c. of the solution in a test-tube. In another test-tube take 5 c. c. of grape-sugar (a raisin soaked in water will give you grape-sugar). To each of these test-tubes add two 2 c. c. of * Fehling's solution of copper sulphate. Heat gently over the burner. Which one gives a yellow to brown precipitate of copper hydrate? Add a few drops of H_2SO_4 to the granulated sugar (cane-sugar); boil. Does it give the yellowish precipitate? This is a delicate test for grape-sugar. Boiling converts the yellow hydrate into a dark-brown sub-oxide. What must be the action of H_2SO_4 upon cane-sugar ($C_{12}H_{22}O_{11}$) in order to change it to $2C_6H_{12}O_6$? Apply the copper test to the following to determine whether the sugar is cane- or grape-sugar: Honey, peach, or any other fruit-juice, molasses, various samples of candy. Try prolonged boiling, and see if that will change cane- to grape-sugar. Put yeast into a solution of cane-sugar and allow to stand twenty-four hours. Has it changed to grape-sugar?

NOTE TO TEACHER. — Try Experiment 134, page 180, Eliot and Storer's "Elementary Chemistry," and test for sugar with Fehling's solution. Try to convert cloth or wood into sugar in the same way.

* Page 140.

Experiment 94.—Digestive Ferments, Albuminous.

NOTE.—The digestive ferments, pepsin, of the gastric juice, pancreatin, of the pancreas, and amylopsin, one of the three ferments of pancreatin, may be obtained from the druggist. They are prepared in powder and may be kept any length of time. By the following experiments try to determine which of the ferments acts upon each of the classes of food: which acts most rapidly; what effect heat and cold have upon digestion; etc. The experiments may be performed by each pupil, or the teacher may perform the experiments before the class, the pupils taking notes as directed.

Boil an average-sized egg ten or fifteen minutes. Separate the yolk from the white and rub each separately through a thirty-mesh sieve. Divide each into three portions by weight and put them into beakers or into two-ounce flat bottles. Number the bottles containing the white 1, 2, and 3 respectively, and those containing the yolk, 4, 5, and 6. Add to each of these 25 c. c. of one-per-cent. **HCl**. Add to Nos. 1 and 4 one centigram or more each of pepsin; to Nos. 2 and 5 one centigram each of pancreatin; and to Nos. 3 and 6 one centigram each of amylopsin. Shake thoroughly and set in a pan of water kept at a temperature of 120° to 130° F. Agitate every five minutes. In which does the albumin dissolve? In which does the oily yolk dissolve? When you are satisfied as to which are not changed by the digestive ferment, neutralize these fluids with sodium carbonate and replace in the water-bath. After sufficient time has elapsed examine again. In which cases now does the food dissolve in the neutralized fluid? Is the gastric juice acid or alkaline? Are the intestinal fluids the same? What class of food is digested in the stomach? What is the temperature of the human body?

Experiment 95.—Effect of Temperature upon Digestion.

Prepare an egg and apparatus as in the last experiment, except that only the white, the **HCl**, and pepsin will be required. Place No. 1 in ice-water, No. 2 in water at 100° to 105° F., and No. 3 in water at 130° F. In which does the albumin dissolve most rapidly? In which most slowly? What effect upon digestion does hot tea, coffee, etc., have, at least so far as temperature is concerned? How would hot food compare with cold food? How would the drinking of ice-water affect digestion for a time?

Experiment 96.—Effect of Mastication upon Albuminous Food.

Prepare an egg as before, except that one third of it should be in a solid chunk; a second, cut into small pieces; a third, passed through the finest sieve. Be careful to use the same quantity of albumin and pepsin and acid for each, and keep them all at the same temperature in a water-bath. In which does digestion take place most rapidly? In which most slowly? Why? Which would resemble rapid eating?

Experiment 97.—Digestion of Starches.

Prepare * iodine solution and Fehling's fluid. Make starch-paste by boiling about 30 grams of laundry starch in about 500 c. c. of water.

Boil until it becomes translucent, stirring thoroughly. Put about 100 c. c. of this paste into each of five bottles, and stand four of them in the water-bath, to be kept at a temperature of 105° to 107° F.

Number the bottles 1, 2, 3, 4, and 5 respectively. Place No. 5 in ice-water. Weigh out carefully one decigram of pepsin for No. 1, one decigram of pancreatin for No. 2, one decigram of amylopsin each for 3 and 5.

Put these weighed quantities of the ferments into their respective bottles, and shake thoroughly. Notice that 4 is only for comparison.

How does the ferment in any case affect the starch mass? In which does it liquefy most? Take five ordinary test-tubes and fill with water. Set them in a row, or number with a gummed label. Drop one small drop of iodine solution into each by means of a pipette, and shake to mix.

Now, when the ferments have been at work for five minutes, take a clean pipette and drop one drop into iodine solution No. 1. Rinse the pipette and do the same with each number successively. Then shake the iodine solution.

No. 4 will show what pure starch and iodine will give. How do the others compare? In which do you find the least change of color by the iodine? Allow to stand five minutes more, and repeat the same, having replenished your test-tubes with iodine solution. What is the effect of temperature as shown by the ice-water?

After you have settled the question as to which of these three ferments is most active in digesting starch, test the product of digestion to determine what the starch has been changed into. Take about 5 c. c. of Fehling's solution, freshly made. Boil it in a test-tube, to be sure it is good. It should not change color.

Now add to it a few drops of digested starch, and boil. What does this show it to be? Test each of the bottles in the same way.

In which is there no, or very little, change, and in which most? Does this agree with the starch test? Where in the alimentary canal is this change made?

Chew a cracker for a few minutes. Does it become sweeter? What does this show? Does the result with regard to temperature agree with the last?

Cane-sugar is converted into grape-sugar before being absorbed, but it is not known by what means it is changed.

* Page 140.

Experiment 98.—Effect of Acids and Alkalies upon Starch Digestion.

Take four bottles of starch, prepared as above, and make 1 and 2 acid with one-per-cent. **HCl**, and 3 and 4 alkaline by $\mathbf{Na_2CO_3}$. Number and add the same portion as above of pepsin to 1 and 3, and amylopsin to 2 and 4. Keep all at the temperature of 105° F., and test as before. What do you learn as to the effects of acids and alkalies upon these ferments in their action toward starches?

Experiment 99.—Digestion of Milk.

Prepare the water-bath, and heat the water to 105° F. Put 25 c. c. of fresh milk in each of the three test-tubes, and leave them in the bath long enough to become heated to that temperature. To 3 c. c. of water add 3 decigrams of sodium carbonate and 1 decigram of pancreatin, and mix thoroughly.

Put this into No. 1, shake, and replace in the bath. Add to No. 2 a like preparation of pepsin and sodium carbonate, and to No. 3, 3 c. c. of one-per-cent. **HCl** and 1 decigram of pepsin. We thus have the acid ferment of the stomach, No. 3, the same neutralized, No 2, and the natural alkaline ferment of the intestine.

How do they appear in the course of five minutes?

Which one has formed a clot or curd?

Which one seems to dissolve up the milk?

To test it, take 5 c. c. of pure milk, and add a drop of dilute $\mathbf{HNO_3}$, and it forms a clot. Do the same with No. 1. Is the milk changed? Is there a clot or a precipitate formed by $\mathbf{HNO_3}$?

Allow to stand for an hour. How have they progressed now? Test for sugar by Fehling's fluid. In which has the milk been prepared most for absorption? Does acid or alkaline pepsin digest milk most rapidly?

If milk forms such a hard clot in the stomach, what could be done to make it digest more quickly? Would lime-water taken with it prevent this? Human milk does not form such a clot as this in an infant's stomach, but remains in small threads.

By the addition of pancreatin to the cow's milk a few minutes before taking, it may be kept from forming hard curds. Is the coagulation of milk due to the ferment or to the **HCl** of the stomach? Try milk and **HCl**. Which of the two ferments digests milk best? What is cheese? How is rennet obtained?

Classifications of Elements and Compounds.—There are about twenty-three elements that act as bases to form the most common salts, and since by far the greatest number of compounds are salts, it is important to know the characteristics of each base, so that we may determine the composition of an unknown compound. If we can perform an experiment by means of some reagent by which the character of the base and the acid radical are brought out, we may then apply the laws of combination, and say of what the salt is composed.

These characteristics are very largely the difference in solubility, color, odor, etc.

In all the following experiments notice carefully these differences and resemblances, and make full notes, so that you may refer to them again and gather up groups of facts as directed. Write out all equations in full, referring back to the Table of Elements, page 28, if you do not know the atomicity of an element, and to the Table of Basicity of Acids, page 34, if you do not know the combining power of an acid. Remember that all that you take must be in the first member of the equation, and all you get in the second member, and that the second member cannot contain an element or a different quantity of the element that did not appear in the first member. In other words, *you cannot make or destroy an atom of matter.*

Experiment 100.—Sulphides.

NOTE TO TEACHER.—Prepare $\mathbf{H_2S}$ upon a large scale, and keep the generating flask in a gas-hood or window where the escaping gas will be carried off. Provide the flask with a delivery-tube to conduct the gas to the bottom of a test-tube. Keep a beaker of water beside the flask, so that the delivery-tube may be washed after each experiment, to prevent mistakes.

Make a test-tube full of the solution of each base to be tested. First try cold water; then, if no solution is formed, boil. If none still, try diluted $\mathbf{HCl}$ and if heat will not bring about the desired end, add $\mathbf{HNO_3}$. If it still refuses to dissolve, heat to fusion on charcoal with $\mathbf{Na_2CO_3}$, and then it will dissolve in acids.

We wish now to see what sulphides we can form by precipitation from a solution of each of the twenty-three bases. Some of them form precipitates with $\mathbf{H_2S}$ from an acid solution, and many will not. Of those that will not thus precipitate sulphides a few *will* if the solution is neutral or alkaline, and since $\mathbf{H_2S}$ is an acid, we must use a soluble salt of that acid, viz., $\mathbf{(NH_4)_2S}$.

To determine these sulphides, divide your test-tube of the solution into two portions, and if the salt was dissolved in water it is probably neutral, since most salts are neutral. Add to one of these portions of the solution a few drops of the acid of the same kind as indicated in the radical of the salt to acidify it. Thus, if the salt is $\mathbf{CuSO_4}$, acidify by a few drops of $\mathbf{H_2SO_4}$. If $\mathbf{NaCl}$, acidify with $\mathbf{HCl}$, etc. If the solution is already acid, neutralize one portion with $\mathbf{NH_4HO}$. Complete the diagram upon the note page, and try each of the twenty-three common bases, using the element or any salt of it.

For a reason to be given later, try to dissolve each of the sulphides obtained with $\mathbf{H_2S}$ from an acid solution, with $\mathbf{(NH_4)_2S}$. The precipi-

tates obtained with aluminum and chromium will be hydrates instead of sulphides; but this fact may be disregarded here. Try each of the bases as follows:

SOLVENT.	SALT OR METAL.	CONDITION.	REAGENT.	PRECIPITATE AND COLOR.	EFFECT OF $(NH_4)_2S$.	REMARK.
H_2O	$CuSO_4$	Neut.	$(NH_4)_2S$		—— ——	
H_2O	$CuSO_4$	Ac.	H_2S	Black ppt.	Insoluble in $(NH_4)_2S$	
HCl	$SbCl_3$	Ac.	H_2S	Orange "	Soluble in $(NH_4)_2S$	
HCl	$SbCl_3$	Neut.	$(NH_4)_2S$			
H_2O	$FeSO_4$	Neut.	$(NH_4)_2S$	Black "		
H_2O	$FeSO_4$	A.	H_2S			

Experiment 101.—Chlorides.

Prepare a table similar to that for sulphides, and try each of the bases to learn the same facts as to what bases form chlorides in an insoluble condition, or a precipitate.

You will notice that although $\mathbf{NH_4HO}$ and **HCl** were united as in Experiment 25 to form a chloride, yet as the **NaCl** is soluble in water, no precipitate appeared. By evaporating to dryness the salt was obtained.

Now, when you put two solutions together there may be a trading of elements, and the new salt formed being soluble either in the water of the solution or the acid present, does not appear as a precipitate.

A chemical change must be detected by some other means than by the formation of a precipitate.

Again, there may be no chemical change, as two substances may not act upon each other at all.

Usually the evolution of a gas or liberating of heat will show that a chemical change is taking place.

Of course you will not expect to precipitate a chloride, by **HCl** from a chloride, so you must use some other acid, as $\mathbf{HNO_3}$, to form a solution if it cannot be done by water, and some other than **HCl** must be used for acidifying.

Do not be discouraged if the number of precipitates is small, for there are but few chlorides that are insoluble in water.

Experiment 102.—Carbonates.

Make solutions as directed, except that it is not necessary to make acid solutions, since the reagent by which the carbonates are obtained, is a strong alkali, and would therefore neutralize the solution.

Since $\mathbf{H_2CO_3}$ (carbonic acid) cannot be made, we must obtain carbonates from a solution of a salt of the acid, as $\mathbf{(NH_4)_2CO_3}$.

As a matter of economy, it is best to neutralize acid solutions before adding ammonium carbonate, and thus save a large quantity of the more expensive carbonate.

Add $\mathbf{HN_4HO}$ to neutralize, and then $\mathbf{(HN_4)_2CO_3}$.

Make a table as before. Try to dissolve the few precipitates formed in $\mathbf{NH_4Cl}$.

It is well to warm the solution after adding the ammonium carbonate if a precipitate does not form readily.

The precipitates obtained will all be carbonates.

Solvent.	Salt or Metal.	Condition.	Reagent.	Precipitate and Color.	Effect of	Remark.

Precipitates obtained by $(NH_4)_2CO_3$.

Solvent.	Salt or Metal.	Precipitate and Color.	Effect of NH_4Cl upon Precipitate.	Remark.
H_2O	$MgSO_4$	Milky-white ppt.	Ppt. dissolves.	Formed slowly only after heating.
H_2O	$AgNO_3$	No precipitate.		

Experiment 103.—Blow-pipe Tests. Metals.

By means of a blow-pipe, a piece of charcoal, and some dry sodium carbonate, and a gas-lamp or alcohol-burner, determine which of the substances indicated give metallic globules, which form a coating upon the charcoal, which are infusible, which give off vapors while heated, what the vapor smells like, which volatilize, and in fact all the changes you notice.

Salts of **Ag, Pb, Hg, As, Sn, Sb, Cu, Ni, Fe, Mg, Zn.**

Experiment 104.—Blow-pipe Test. Salts.

In a similar way to the last experiment determine which of the salts named deflagrate and burn the charcoal, which fluff up and form a charred mass, which give off the odor of sulphur, which give off water, as shown by fusing and then becoming dry, etc. No sodium carbonate need be used in this test; simply the charcoal, plow-pipe, and lamp. Try chlorates, nitrates, sulphates, oxalates, tartrates, sulphides.

Experiment 105.—Borax-bead Test.

Make a small hook on the platinum wire the size of the letter *o* in the word borax in the head-line, and heat it; thrust it into the borax-bottle and heat till you have a clear bead.

Now take a small quantity of the dry powder and touch the hot bead to it. Heat it again until it becomes clear. Note the color. Make a new bead for each. Strike the bead a sharp blow, and it will crumble out of the loop.

Which of the salts named below give some decided color to the bead? What is the color of each? $FeSO_4$, $Ni(NO_3)_2$, $CuSO_4$, $NaCl$, $CO(NO_3)_2$, MnO_2, $K_2Cr_2O_7$.

Experiment 106.—Flame Test.

Pulverize the salts to a fine powder. Moisten the platinum wire in **HCl**, dip it into the powder, and burn in the flame of the lamp.

Clean the wire thoroughly after each test, either by burning it until it gives no color or by rubbing it with emery and then burning it a little.

Note the color of the flame of each of the following:

$CuSO_4$, $BaCl_2$, $Sr(NO_3)_2$, $CaCl_2$, KCl, Na_2CO_3.

Now try other salts of these metals, and see if the color is due to the base or to the acid radical.

GROUPS OF BASES.

If the work has been carefully and properly done, you should have found that but three bases form chlorides as insoluble precipitates in water, viz., **Pb**, **Ag**, and **Hg**.

We may use this means of separating these three bases from the others, and for convenience we will call this the First Group, or the **HCl** Group.

The acid solution of bases should have given precipitates with **As**, **Sb**, **Sn**, **B**, **Cu**, **Cd**, as well as the First Group.

Let us discard the three bases already distinguished from the rest by their chlorides, and we have a Second Group, the $\mathbf{H_2S}$ Group. Three of these dissolved upon the addition of $\mathbf{(NH_4)_2S}$, viz., **As**, **Sb**, and **Sn**, and the other three did not.

We may designate the soluble sulphides as Division A of the $\mathbf{H_2S}$ Group, and the others as Division B.

From the neutralized solution and $\mathbf{(NH_4)_2S}$, were formed the following: **Fe**, **Co**, **Ni**, **Zn**, and **Mn**, sulphides, and by the action of the reagent **Al** and **Cr** were precipitated as hydrates.

Classing all these together, we have a Third Group, the $\mathbf{(NH_4)_2S}$ Group.

Ammonium carbonate should give precipitates of four carbonates, **Ba**, **Sr**, **Ca**, and **Mg**, the latter only being soluble in $\mathbf{NH_4Cl}$.

Call these four the Fourth Group, or $\mathbf{(NH_4)_2CO_3}$ Group.

Of the twenty-three common bases there are left, the two elements **K** and **Na**, and the radical $\mathbf{NH_4}$, which forms a common base.

These three may be shown to be closely related, and are therefore classed as the Fifth Group, or Alkaline Group.

Their salts, so far as known, are all readily soluble in water (see Table I., page 107). They may be put to similar uses, as, for example, bicarbonate of soda, or bakers' soda, bicarbonate of potash, or saleratus, and bicarbonate of ammonium are all used to raise biscuit.

The chlorides of each differ in physical properties, chiefly as regards permanence in exposure to air and moisture. The hydrates may be used for the same purposes in softening water, neutralizing acids, etc.

Before Sir Humphry Davy isolated the elements **K** and **Na**, they were called the vegetable alkali, **K**, the mineral alkali, **Na**, and the volatile alkali, ammonium. (For the analysis of an unknown base, see Key, page 112.)

These groups are only arbitrary divisions, based upon the resemblance with reference to one or two reagents, called group reagents.

TABLE I.—TABLE SHOWING SOLUBILITY OF THE COMMON SALTS, OXIDES, AND HYDRATES IN WATER AND ACIDS.

	Acetates.	Arsenates.	Arsenites.	Borates.	Bromides.	Carbonates.	Chlorates.	Chlorides.	Chromates.	Cyanides.	Ferricyanides.	Ferrocyanides.	Fluorides.	Hydrates.	Iodides.	Nitrates.	Oxalates.	Oxides.	Phosphates.	Silicates.	Sulphates.	Sulphides.	Tartrates.
Al	W	A		A	W		W	W					W	A	W	W	A	A	A	sa	W	A	W
NH_4	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W		W		W	W	W
Sb	W	A	A		sw			sw	A				W		sw		A	A	sw		sw	A	A
Ba	W	A	A	A	W	A	W	W	A	sw	sw	sw	sa	W	W	W	A	W	A	A	I	W	A
Bi	W	A		A	sw	A	W	sw	A	sw		I	W	A	A	W	A	A	A		W	A	A
Cd	W	A		sw	W	A	W	W	A	A	A		sw	A	W	W	A	A	A	A	W	A	sw
Ca	W	A	A	A	W	A	W	W	sw	W	W	W	sa	sw	W	W	A	sw	A	A	sa	W	A
Cr	W	A		A	W		W	W	A	A			W	A	W	W	sw	A	A	A	W	A	W
Co	W	A	A	A	W	A	W	W	A	sa	I	I	sw	A	W	W	A	A	A	A	W	A	W
Cu	W	A	A	A	W	A	W	W	W	A	I	I	A	A	W	W	A	A	A	A	W	A	W
Au					W			W		A				A	A			A				A	
H	W	W		W	W		W	W		W	W	W	W	W	W	W	W	W	W	W	W	sw	W
Fe''	W	A	A	A	W	A	W	W		sa	I	I	sw	A	W	W	A	A	A	A	W	A	sw
Fe'''	W	A	A	A	W		W	W	W		W	I	W	A	W	W	A	A	A	A	W	A	W
Pb	W	A	A	A	sa	A	W	sa	sa	A	sw	A	A	A	sw	W	A	A	A	A	sa	A	A
Mg	W	A	A	sw	W	A	W	W	W	W	W	W	sa	A	W	W	A	A	A	A	W	A	W
Mn	W	A	A	A	W	A	W	W	W	A	I	A	A	A	W	W	sw	A	A	A	W	A	sw
Hg''	sw	A	A		sa	A	W	sa	A	W			W		A	W	A	A	A		sw	A	sw
Hg'	W	A	A		W	A	W	W	sw	W	A	A	sw		A	W	A	A	A		W	A	A
Ni	W	A	A	A	W	A	W	W	A	sa	I	I	sw	A	W	W	A	A	A	A	W	A	A
K	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W
Ag	sw	A	A	A	I	A	W	I	A	I	I	I	W		I	W	A	A	A		sw	A	A
Na	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W
Sn''''	W	A			W			W				I	W	A	W	sw	W	A	A		W	A	
Sn''	W	A	A	A	W		W	W	A		I	I	W	A	W		A	A	A		W	A	A
Sr	W	A	A	A	W	A	W	W	sw	W	W	W	sa	A	W	W	A	sw	A	A	sa	W	A
Zn	W	A		A	W	A	W	W	W	A	A	sa	sw	A	W	W	A	A	A	A	W	A	A

W = soluble in water; A = soluble in acids; sw = sparingly soluble in water but soluble in acids; sa = sparingly soluble in acids; I = insoluble in acids.

FIRST GROUP METALS.

Pb, Ag, Hg′.

SEPARATION AND IDENTIFICATION OF METALS OF THE FIRST GROUP.

Experiment 107.

(*a*) Measure out 5 c. c. $\mathbf{Pb(NO_3)_2}$ solution in your graduated test-tube and add a few drops of **HCl**. What salt is formed?

(*b*) Filter and wash the precipitate with a little cold water.

(*c*) Pierce the bottom of the filter-paper with your platinum wire, and wash a quantity of the precipitate through into a clean test-tube.

(*d*) Boil the precipitate in water, filter, and test the filtrate by adding a few drops of $\mathbf{K_2CrO_4}$, and chrome yellow will be precipitated.

Experiment 108.

Repeat the above, using $\mathbf{AgNO_3}$ solution in place of $\mathbf{Pb(NO_3)_2}$ solution. Is $\mathbf{PbCl_2}$ dissolved by hot water? Is **AgCl** dissolved in hot water? Thus we have a means of separating lead and silver in solutions of their salts.

Experiment 109.

(*a*) Make a small quantity of mercurous chloride and try to dissolve it in hot water. Does it dissolve?

(*b*) Try to dissolve a small quantity of **AgCl** in $\mathbf{NH_4OH}$ by placing in test-tube and heating.

Try same experiment with $\mathbf{Hg_2Cl_2}$. Does it dissolve in $\mathbf{NH_4OH}$?

Thus we have a method of separating **Pb**, **Ag**, and **Hg** (ous) in solutions of their salts.

Experiment 110.

Take 5 c. c. solution containing all three salts, $\mathbf{Pb(NO_3)_2}$, $\mathbf{AgNO_3}$, and $\mathbf{HgNO_3}$, and add **HCl**. Filter, wash precipitate thoroughly in cold water, and throw away the filtrate.

(*a*) Add 10 or 15 c. c. boiling water to the precipitate on filter-paper, and test filtrate with few drops of $\mathbf{K_2CrO_4}$. Which salt is dissolved by the hot water and precipitated upon addition of $\mathbf{K_2CrO_4}$? Write the equation for the last reaction.

(*b*) To the residue in the filter add $\mathbf{NH_4OH}$. Which salt is dissolved? Which one is not dissolved? Test filtrate by adding few drops of nitric acid. The equation is:

$$\mathbf{(NH_3)_3(AgCl)_2 + 3HNO_3 = 2AgCl + 3NH_4NO_3}.$$

(*c*) What salt is blackened upon filter-paper?

Its formula being $\mathbf{NH_2Hg_2Cl}$, write equation.

Experiment 111.—Confirmatory Tests for First Group.

Take 25 c. c. of each of the solutions of these first-group metals, and test each by adding a few drops of each of the following reagents:

KOH or **NaOH** in excess.

NH_4OH, H_2S, H_2SO_4, **KBr**, **KI**, and K_2CrO_4.

Write equation in each case, and underline precipitate; as,

$$AgNO_3 + HCl = \underline{AgCl} + HNO_3.$$

Compare precipitates when formed, as to color, solubility in HNO_3 and NH_4OH, and tabulate your results.

Experiment 112.—Reduction of Silver.

Make a solution of common salt (**NaCl**) in a test-tube. Take 1 c. c. of $AgNO_3$ in another test-tube, and add to it the salt solution until all the silver has been precipitated. $AgNO_3 + NaCl = \underline{AgCl} + ?$ Filter and transfer to a piece of charcoal. Mix with the precipitate on the charcoal as much sodium carbonate, Na_2CO_3. Heat before the blow-pipe until you get a bead of metallic silver. How can **NaCl** form **AgCl** as well as **HCl** does? How do sodium carbonate and charcoal take away chlorine from the silver?

Test the silver bead with H_2SO_4, **HCl**, and HNO_3, to ascertain which will dissolve it most readily. Put the bead in a test-tube, and add a few drops of acid, and heat. As soon as you are satisfied as to whether chemical action takes place or not, pour in water and rinse. In the same way try the other acids. Which is the best solvent of silver?

Can you reduce lead or mercury in the same way?

Optional Experiment 113.—Pure Silver from Silver Coin.

Put a small silver coin into a beaker and cover it with HNO_3 and water, about equal parts. Heat gently until the coin is dissolved. The blue color is due to copper. Why is copper used in silver coins? Add **HCl** until no more silver can be precipitated. Filter, and wash by pouring water over it until the liquid runs through colorless. Transfer the silver chloride to a porcelain dish, add some clippings or grains of zinc, and cover with dilute H_2SO_4. What gas will be generated by zinc and sulphuric acid? As this gas escapes it finds an element, **Cl**, for which it has strong affinity. It unites with **Cl** to form **HCl**, leaving the silver free. In an hour the silver will all be set free, and appear as a black, spongy mass. It may be collected and melted into a globule on charcoal, or it may be dissolved in HNO_3 to form silver nitrate. The common name of silver nitrate is lunar caustic. For what is it used?

Experiment 114.—Silver-printing.

Float a small sheet of paper upon a solution of silver nitrate in a dark room. Dry in the dark, and then lay a fern leaf, a piece of lace, or a photographic negative upon the silver paper.

Cover with glass and place in the sunshine until the exposed parts are blackened. Rinse with water in the dark room, and then for a few minutes in a solution of hyposulphite of sodium, to dissolve out all the silver not acted upon by light.

Then rinse for some time in water and dry.

How long did it take to blacken the silver salt in the sunlight?

Could photographic negatives be made by use of this salt?

If not practicable, what salt is used because so quick that it is instantaneous?

Is silver nitrate used at all in photography?

Experiment 115.—Silver-plating.

(*For the Teacher.*)—Let the teacher prepare cyanide of silver and plate a key or some such article to illustrate the process of silver-plating.

See Gore's "Electro-Metallurgy" for process, solutions, etc.

What force separates the silver from cyanogen?

Upon which pole of the battery does silver collect?

Why is it best to keep a piece of metallic silver upon the other pole?

How is this silver affected?

Why does the silver not adhere where there is grease or finger-marks on the article to be plated?

Write a full account of the process.

Experiment 116.—Reduction of Lead by Zinc.

Suspend a piece of metallic zinc in a solution of lead nitrate or acetate. Allow it to stand over night. Zinc will exchange with the lead of the solution. Notice the tree-like growth of gray deposit of lead. Test the solution for zinc as follows:

First, to precipitate the lead not yet exchanged, add alcohol, and then sulphuric acid. Alcohol prevents the sulphate of lead from being dissolved in water as fast as precipitated by the acid. Now filter out the sulphate of lead and neutralize the filtrate with ammonia and then add $(NH_4)_2S$, and the zinc will be precipitated as a white sulphide.

What color is lead sulphide?

Complete the equation $Zn + Pb(NO_3)_2 = ?$

Experiment 117.—Distinction between Mercuric and Mercurous Salts.

Dissolve a globule of mercury in HNO_3, using more of the acid than necessary to dissolve the globule, and you will have mercuric nitrate, $Hg(NO_3)_2$. (Remove gold rings from the fingers while working with mercury.)

Partly dissolve another globule in HNO_3 and pour off the acid into another test-tube, and you will have mercurous nitrate, $Hg_2(NO_3)_2$. Dilute these solutions with water and apply the following tests:

(*a*) Try a little of each with H_2S. What is the result?

(*b*) Try a little of each with **NaHO**.

(*c*) Try a little of each with **HCl**. *Which one gives a precipitate?*

(*d*) Introduce a clean strip of copper or a clean copper coin into each. Rub with cloth after it becomes blackened. Try zinc. Why is the precaution given above in regard to gold rings?

Sum up the differences between mercuric and mercurous compounds.

It will be noticed that mercurous compounds belong to the First Group, and that mercuric compounds belong to the Second Group. Also that lead is not all or completely precipitated by **HCl**; so in the key at the end of the book it is to be found also in the Second Group.

Experiment 118.—Reduction of Mercury.

Put a small fragment of the common ore of mercury **HgS**, called cinnabar, into a hard glass tube open at both ends. Hold the tube inclined upward in the flame, and heat strongly. What gas is given off? What indication of metallic mercury? Why must the tube be open at both ends?

SECOND GROUP METALS.

Sb, Sn, Hg″, Bi, Cu, and Pb.

NOTE.—In testing solutions of unknown substances, the lead, if present, is not all precipitated by **HCl**, and hence must be looked for again in the Second Group.

Experiment 119.—Separation of Second Group into Two Divisions.

Have solutions of the following reagents ready for use: $SbCl_3$ (**HCl** solution), $SnCl_2$, $Hg(NO_3)_2$, $Pb(NO_3)_2$, $Bi(NO_3)_2$, and $Cu(NO_3)_2$. Add dilute **HCl** to all but $PbNO_3$ until each gives acid reaction with litmus.

(*a*) To 5 c. c. of $SbCl_3$ add quantity strong solution H_2S and filter. Wash precipitate thoroughly in quantity of cold H_2O, and throw away the filtrate.

(*b*) Pierce the filter-paper with a platinum wire or a glass rod, and wash the precipitate through into an evaporating dish.

(*c*) Add $(NH_4)_2S_2$, yellow ammonium sulphide, using as little as possible, and stir for some time. The precipitate should dissolve.

Experiment 120.

Perform the same experiments, using $SnCl_2$ instead of $SbCl_3$. Does the precipitate dissolve at point (*c*)?

Experiment 121.

Perform the same experiments, using $Cu(NO_3)_2$. Does precipitate (CuS) dissolve in $(NH_4)_2S_2$?

Experiment 122.

Do the same with $Bi(NO_3)_2$, $Hg(NO_3)_2$, and $Pb(NO_3)_2$. Separate metals of the Second Group into two divisions, based on the solubility of their sulphides in $(NH_4)_2S_2$.

Experiment 123.—Identification of Sb in Solution of its Salts.

Add H_2S to $SbCl_3$ solution. Note the color of the precipitate in $(NH_4)_2S_2$ and re-precipitate by using dilute HCl.

NOTE.—The "spot" test for antimony is one of the best, and may be performed by the teacher if thought desirable. Arsenic also belongs to this group. This and many other elements may be studied with profit in more advanced classes.

Experiment 124.—Identification of Sn.

(*a*) Wash thoroughly on the filter the precipitate obtained by adding H_2S to $SnCl_2$ solution. What is the color of this precipitate?

(*b*) Dissolve in yellow ammonium sulphide and re-precipitate with HCl, as in the case of antimony. What is the color of this precipitate? Is it the same as the color obtained in (*a*)?

Experiment 125.—Sn, Continued.

(*a*) Place some tin salt upon charcoal and use the oxidizing flame. SnO_2 is formed about the edge of the substance. Note its color when hot and when cold. Moisten the oxide when cold with cobaltous nitrate, and heat again. Note color.

(*b*) To solution of $SnCl_2$ add some $HgCl_2$. Color of precipitate? Allow contents of tube to stand, and note change. Do you find any metallic Hg?

(*c*) Decompose solution of $SnCl_2$ by means of electric current.

(*d*) Add NH_4OH to solution of $SnCl_2$ and put in a strip of Zn.

NOTE.—The separation of Sn from Sb may be left for a more advanced course.

IDENTIFICATION OF METALS IN SECOND DIVISION OF SECOND GROUP.

Cu, Bi, Hg″, and Pb.

Experiment 126.—Copper.

(*a*) Dissolve a small bit of copper in nitric acid. What gas is given off? Color of solution? To what salt is this due?

(*b*) Heat a little copper sulphate in evaporating dish. Note change of color as water is driven off. Try to restore color with a drop of alcohol. Try water.

(*c*) Precipitate **CuS** from solution of $\mathbf{Cu(NO_3)_2}$. Wash and dissolve precipitate on the filter-paper with hot $\mathbf{HNO_3}$.

Evaporate the solution thus obtained under the hood in the laboratory or outside the window.

Dissolve the residue in small quantity of water.

If a slight white precipitate is formed in the water at this point, add a drop or two of $\mathbf{HNO_3}$ until it disappears.

Add $\mathbf{NH_4OH}$. Blue solution indicates *copper*.

Experiment 127.—Confirmatory Tests for Copper.

(*a*) Dip a borax bead into a solution of some copper salt, and heat before oxidizing flame. The bead should be green when hot, blue when cold.

(*b*) Place two drops of copper solution on a clean knife-blade. In a few moments dip the blade into water and note copper, reduced from its salt by iron.

(*c*) Allow the clean copper wire to stand for a few minutes in solution of $\mathbf{AgNO_3}$.

What is formed on its surface?

Explain.

(*d*) Place a drop of mercury in a watch-glass and cover with concentrated solution of $\mathbf{AgNO_3}$. Explain the phenomenon.

Experiment 128.—Bismuth.

Dissolve the sulphide formed by adding $\mathbf{H_2S}$ to the solution in $\mathbf{HNO_3}$, and add excess of $\mathbf{NH_4OH}$.

White precipitate is $\mathbf{Bi(OH)_3}$.

(For experiments on **Hg″** and **Pb**, see those metals under First Group.)

To a nearly neutral solution of a bismuth salt add a large quantity of water, and a precipitate will be formed.

Experiment 129.—Separation of Metals in Second Division of Second Group.

(*a*) Dissolve precipitate left after removing **Sb** and **Sn** from a solution containing salts of all the Second Group metals, in hot nitric acid. In case it will not all dissolve, the residue may be mercury ($\mathbf{Hg''}$), which is now in the form of **HgS**, and may be dissolved in nitro-hydrochloric acid. To this solution add $\mathbf{SnCl_2}$. The reaction which occurs is the test for $\mathbf{Hg''}$.

(*b*) Evaporate solution (*a*) to dryness, and dissolve residue in water. If a white precipitate forms, add a drop or two of $\mathbf{HNO_3}$ to clear. To this solution add a drop or two of $\mathbf{H_2SO_4}$. A white precipitate shows that **Pb** is present and $\mathbf{H_2SO_4}$ should be added until all **Pb** is precipitated as $\mathbf{PbSO_4}$. To the filtrate at this point, or in case there is no lead, to solution obtained in (*b*) add excess of $\mathbf{NH_4OH}$. White precipitate denotes **Bi**. Blue colored solution indicates copper.

Write the equation for the reaction when bismuth is present.

THIRD GROUP METALS.

Fe, Cr, Al, Ni, Co, Mn, Zn.

Experiment 130.—Fe.

Dissolve a piece of iron wire, two inches long, in **HCl** in a test-tube.

Write the equation for the reaction.

NOTE.—Iron acts as a base in forming two series of salts, the *ferrous* and the *ferric*. The ferrous salts may be identified as such by adding a few drops of $\mathbf{K_3FeCy_6}$, potassium ferricyanide, which gives a dark-blue precipitate, $\mathbf{Fe_3(FeCy_6)_2}$, ferrous ferricyanide.

Test a portion of the solution just made and dilute with water, and determine whether the salt contained is $\mathbf{Fe_2Cl_6}$ or $\mathbf{FeCl_2}$.

Experiment 131.—Fe.

Dilute a portion of the remaining solution prepared in the preceding experiment, and add two or three drops of strong nitric acid. Is there any change of color?

Boil carefully for a few moments and cool. Test a small portion of the contents of your test-tube with a few drops of $\mathbf{K_3FeCy_6}$.

Do you obtain the same results as before? This shows the salt of iron to be in the ferric condition.

The action of $\mathbf{HNO_3}$ upon **Fe** in changing it from the ferrous ($\mathbf{Fe''}$) to the ferric ($\mathbf{Fe'''}$) condition may be represented in the following equation:

$$2Fe + 8HNO_3 = Fe_2(NO_3)_6 + 2NO + 4H_2O.$$

Experiment 132.—Fe.

Dissolve two or three crystals of $\mathbf{FeSO_4}$ in water in your test-tube.

NOTE.—Ferrous salts, on being exposed to the air, become coated with a layer in the ferric condition. Hence, in getting solution of $\mathbf{FeSO_4}$, after dissolving away a portion from the outside, and when the crystals have become a clear green in color, throw this solution away, add fresh distilled water, and warm gently to hasten solution.

This solution, if used as soon as prepared, is in a ferrous condition, but will become ferric on standing.

Test a portion of the solution with $\mathbf{K_3FeCy_6}$, as before.

Separate the solution into several portions, and test with $\mathbf{K_4FeCy_6}$ and with **KCyS**.

Name these substances.

Oxidize some of the $\mathbf{FeSO_4}$ to $\mathbf{Fe_2(SO_4)_3}$, and test with the three reagents named above.

Make a table showing results obtained with each in ferrous and ferric solutions.

Experiment 133.—Fe.

(*a*) Make a fresh solution of $\mathbf{FeCl_2}$ as in Experiment 130, and test a small portion with $\mathbf{K_4FeCy_6}$, $\mathbf{K_3FeCy_6}$, and **KCyS**.

(*b*) Pass **Cl** gas into the remainder. [This should be done under the hood.]

Test with the above three reagents, and compare results with those tabulated from preceding experiment.

What is the effect of **Cl** upon ferrous salts?

Complete the equation, $\mathbf{2FeCl_2 + 2Cl} = ?$

Experiment 134.—Fe.

Under the hood pass $\mathbf{H_2S}$ gas into dilute solution of $\mathbf{Fe_2Cl_6}$, until a few drops of $\mathbf{K_3FeCy_6}$ reveals the presence of the salt in the ferrous condition.

Complete the equation, $\mathbf{Fe_2Cl_6 + H_2S} = \mathbf{2FeCl_2} + ?$

In the analysis of unknown solutions, what reagent is added to precipitate Second Group metals?

Is some of it left in the filtrate?

How can you tell?

If iron salts are present, what is their condition, ferrous or ferric, when you are ready to add the Third Group reagents?

If you began with a ferric salt, what has produced the change?

Experiment 135.

Add NH_4Cl and NH_4OH to solution of *ferrous* sulphate. Note result.

To solution of ferrous sulphate add a few drops of strong HNO_3, and heat. To this, when cool, add NH_4Cl and NH_4OH as before. Compare results. Explain.

Thus we see that if iron is present in an unknown solution in the form of ferric salts, it is reduced to the ferrous condition by addition of the Second Group reagent H_2S, and must be oxidized to the ferric condition by the addition of HNO_3 and application of heat before it will be precipitated as $Fe(OH)_6$ by NH_4Cl and NH_4OH.

Experiment 136.—Al.

Make a solution of alum. Add NH_4Cl solution and NH_4OH.

This precipitates **Al** as $Al_2(OH)_6$.

Note character of the precipitate and compare with $Fe_2(OH)_6$ formed in the same manner. What is alum?

Balance the equation $Al_2Cl_6 + 6NH_4OH = ?$

Experiment 137.—Cr.

To a solution of chrome alum, $K_2Cr_2(SO_4)_2$, add some NH_4Cl solution and some NH_4OH. **Cr** is precipitated as $Cr_2(OH)_6$. Compare precipitate with those of iron and **Al** produced in the same manner.

Fe, **Al**, and **Cr** are thus seen to resemble each other in this, that they form precipitates (hydrates) upon the addition of NH_4Cl and NH_4OH to solutions of their salts.

NOTE *a.*—As has been shown, iron must be in the ferric condition.

NOTE *b.*—**Cr** will not be precipitated as a hydrate at this point, unless it is present as a base. To insure this, **HCl** and H_2S must be added to solutions which we are analyzing, even if there are no First or Second Group metals present.

NOTE *c.*—**Cr** may be reduced to the condition of a base in still another manner, if preferred. In case solution with which you begin is $K_2Cr_2O_4$, add a few drops of **HCl** and a few drops of C_2H_6O, boil and cool. This gives Cr_2Cl_6.

1. Hence, in examining solutions for **Fe**, **Cr**, and **Al**, we add **HCl** and H_2S to our solution to remove First and Second Group metals.

2. To the filtrate, if any are present, or to this solution, we add a few drops of strong HNO_3, boil until no odor of H_2S is given off, and cool.

3. When cold, we are ready for the Third Group reagents, which are NH_4Cl, NH_4OH, and $(NH_4)_2S$.

4. As has been seen, the addition of the first two serves to separate **Fe**, **Al**, and **Cr** from the remainder of the Third Group, **Ni**, **Co**, and **Mn**, which will be precipitated as sulphides upon the addition of $(NH_4)_2S$.

Experiment 138.—To Separate Metals of the Third Group into Two Divisions.

Treat a solution containing salts of all three metals of this group as suggested for iron and chromium. (See 1 and 2 above.)

Experiment 139.

Add $\mathbf{NH_4Cl}$ and $\mathbf{NH_4OH}$, and filter. Test filtrate to see that all of the **Al**, **Fe**, and **Cr** is precipitated. Filtrate contains **Ni**, **Co**, **Mn**, and **Zn**. To this add $\mathbf{(NH_4)_2S}$ and obtain **NiS**, **CoS**, **MnS**, and **ZnS**.

Experiment 140.—To Separate and Identify Fe, Al, and Cr.

Using precipitate obtained on filter in preceding experiment, wash through into glass beaker or large test-tube, add quite a quantity of **KOH**, and boil for some time.

The **KOH** dissolves the $\mathbf{Al_2(OH)_6}$. Set aside this solution to test for **Al**. The remaining precipitate, if there is any, consists of $\mathbf{Fe_2(OH)_6}$ or $\mathbf{Cr_2(OH)_6}$, or both.

Experiment 141.

Separate this precipitate into two portions and dissolve one in **HCl**, and employ test for iron given in Experiment 132. Fuse the other portion of the precipitate on platinum-foil with sodium or potassium nitrate and carbonate.

Dissolve the mass obtained ($\mathbf{Na_2CrO_4}$) in water, and apply tests for chromium as shown in the following experiment:

Experiment 142.—Cr Confirmatory Tests.

To a portion of the solution of $\mathbf{Na_2CrO_4}$ obtained in the preceding experiment, add acetic acid until the solution gives an acid reaction with litmus.

To this add lead acetate solution.

Note color of precipitate ($\mathbf{PbCrO_4}$).

Write equation.

Experiment 143.

Borax bead is given an emerald-green color.

Experiment 144.—Al.

Acidulate the potassium solution of **Al** obtained in Experiment 140 with **HCl**. To this add $\mathbf{(NH_4)_2CO_3}$.

Precipitate $\mathbf{Al_2(OH)_6}$ indicates **Al**.

Experiment 145.—To Separate and Identify Ni, Co, Mn, and Zn.

After removing metals of the first division of the Third Group, add $(NH_4)_2S$, and sulphides of **Ni**, **Co**, **Mn**, and **Zn** are precipitated.

Filter, and save filtrate to examine for Fourth Group metals.

Treat precipitate on the filter with dilute **HCl**.

The black residue is **Co** or **Ni**, or both, and filtrate contains **Zn** or **Mn**, or both.

Experiment 146.

Test some of the residue with borax bead and blow-pipe.

A blue bead indicates **Co**.

If **Ni** is present alone, the bead will have a brown color when hot, and yellow when cold, in the oxidizing flame only.

Experiment 147.

Dissolve the residue in nitro-hydrochloric acid.

If **Ni** is present alone, $Ni(OH)_2$ will be precipitated upon the addition of a few drops of ammonia.

The color of the precipitate is green.

Experiment 148.

If **Co** and **Ni** are both present, the tests interfere with each other to some extent, and the method of separating them is omitted.

Experiment 149.

The filtrate from **Co** and **Ni**, which may contain **Mn** or **Zn**, or both, contains them in the form of $ZnCl_2$ and $MnCl_2$.

Boil the filtrate to expel H_2S, then cool it, and add a decided excess of **KOH**.

Allow to stand for some time. If **Mn** is present, it will be precipitated as $Mn(OH)_2$.

Experiment 150.

With acetic acid acidulate the filtrate from this. Add $(NH_4)_2S$, and **ZnS** is precipitated.

ZnS is insoluble in acetic acid.

Note the color of the precipitate.

How does it differ from the other sulphides of this division of the group?

Experiment 151.—Fourth Group Metals.

Place a small piece of freshly burned quick-lime in a large evaporating dish, and pour on water until the lime is nearly covered.

Note the action and describe it.

After the action ceases, is there as much water present as before? What has become of it? What substance is formed? What is mortar?

Experiment 152.

Try $\mathbf{CaCO_3}$, $\mathbf{Ca(OH)_2}$, and $\mathbf{CaCl_2}$ as to their solubility in water.

In making lime-water for experiments with $\mathbf{CO_2}$, what substance was dissolved? What substance is formed when $\mathbf{CO_2}$ is passed into lime-water? Is it soluble? How do you know?

Experiment 153.

Heat moderately some powdered gypsum ($\mathbf{CaSO_4 2H_2O}$) in a hard glass test-tube or evaporating dish. What happens? See if the powder which is left will harden when made into a paste by the addition of a little water. Explain the action. What is "plaster of Paris"?

Experiment 154.

Fasten a small quantity of asbestos in a loop of platinum wire. Dip into a solution of $\mathbf{CaCl_2}$ and heat in the non-luminous Bunsen flame in front of the spectroscope. Make a drawing of the spectrum in your note-book. Is the spectrum continuous or one of the "bright-line" variety?

Compare also spectrums of $\mathbf{SrCl_2}$ and $\mathbf{BaCl_2}$. What color is imparted to the flame in each instance?

Experiment 155.

Make a solution of $\mathbf{CaCl_2}$ by dissolving a small piece of marble ($\mathbf{CaCO_3}$) in **HCl**.

Write the equation.

Add $\mathbf{NH_4OH}$, $\mathbf{NH_4Cl}$, and $\mathbf{(NH_4)_2CO_3}$.

The substance formed is a carbonate.

Write the equation.

Thus we see that **Ca** is precipitated as $\mathbf{CaCO_3}$ in the presence of $\mathbf{NH_4Cl}$ and $\mathbf{NH_4OH}$. The chlorides, hydroxides, and sulphides of **Ca** are not precipitated when **HCl**, $\mathbf{NH_4OH}$, or $\mathbf{(NH_4)_2S}$ are added to precipitate metals of the first three groups. Hence this serves to form the basis of a new group, the fourth.

Ba and **Sr** are precipitated in the same way.

Experiment 156.

To the filtrate from the Third Group (Experiment 145) or to a solution containing no First, Second, or Third Group metals, add $\mathbf{NH_4OH}$, $\mathbf{NH_4Cl}$, and $\mathbf{(NH_4)_2CO_3}$.

Precipitate indicates **Ba**, **Ca**, **Sr**, or any two of them, or all (see Experiment 147). They are precipitated as carbonates.

Write the equations, using the solutions of $\mathbf{BaCl_2}$, $\mathbf{SrCl_2}$, and $\mathbf{CaCl_2}$.

Save the precipitate on the filter for use in Experiment 158. Save filtrate for Experiment 160.

Experiment 157.

(*a*) Add some of each of the Fourth Group reagents to solution of $\mathbf{BaCl_2}$.

Filter, wash precipitate, and try to dissolve it in dilute acetic acid.

(*b*) Try same with solution of $\mathbf{CaCl_2}$.

(*c*) Try same with solution of $\mathbf{SrCl_2}$.

Experiment 158.

(*a*) Dissolve precipitate on filter in Experiment 156. Substances in solution at this point are acetates of one or more of the Fourth Group metals.

Write the equations.

(*b*) To a small portion of the solution add some $\mathbf{K_2Cr_2O_7}$ solution.

Precipitate denotes barium. Yellow precipitate is $\mathbf{BaCrO_4}$.

If barium is present, add $\mathbf{K_2Cr_2O_7}$ to the rest of the solution, and remove it all. Filter and test filtrate for **Ca** and **Sr** as follows:

(*c*) Precipitate by $\mathbf{(NH_4)_2CO_3}$ and $\mathbf{NH_4OH}$. Filter and wash precipitate.

Dissolve it in **HCl**. Evaporate (under hood) nearly to dryness, and make concentrated solution by adding a very little water. Divide solution into two portions.

1. Test one for **Sr** by adding solution of calcium sulphate and boiling. Set aside to cool, and in about fifteen minutes a precipitate denotes the presence of **Sr**.

2. Test the second portion by adding a solution of $\mathbf{K^2SO^4}$ and filtering, to make certain that there is no **Sr** present. To the filtrate add $\mathbf{NH_4OH}$ and $\mathbf{(NH_4)_2C_2O_4}$ and precipitate denotes **Ca**.

Metals whose chlorides, sulphides, and carbonates are soluble, and hence not precipitated by any of the group reagents used so far, comprise the Fifth Group.

THE FIFTH GROUP.

Mg, K, Na, and NH_3.

Experiment 159.—Magnesium.

To the filtrate from Group Four (Experiment 156), or an original solution, add $\mathbf{Na_2HPO_4}$. Precipitate indicates **Mg**.

It is better to add a little $\mathbf{NH_4OH}$, and then $\mathbf{NH_4Cl}$, before adding $\mathbf{Na_2HPO_4}$, when a fresh solution of $\mathbf{MgCl_2}$ is being tried for the test.

Experiment 160.—Ammonia.

Add **KOH** or **NaOH** solutions to your solution, and warm gently. Odor of $\mathbf{NH_3}$ is a test. Note effect of vapor of $\mathbf{NH_3}$ on moistened litmus paper. Hold glass rod moistened with **HCl** near the tube containing the solution.

What is formed on the rod?

Write the equation.

Experiment 161.—Sodium.

Sodium is detected in solution of its salts by the color which they give to the (*a*) non-luminous Bunsen flame, and by the spectrum. Note these carefully. See experiments on **Ba**, **Ca**, and **Sr**, and compare spectra.

(*b*) Review your experiments in making common salt and in decomposing water by the action of metallic potassium. Will metallic **Na** do as well?

Experiment 162.—Potassium, K.

Examine color given to flame, as in Experiment 151, when **KCl** solution is used. Examine its spectrum.

These are tests of **K** in solution of its salts.

Experiment 163.

Examine the metallic **K** and **Na** kept in the laboratory. Why are they kept under oil? What element must be kept under water?

Experiment 164.

Place some wood ashes on a filter and pour on water. As fast as it runs through pour it back, over and over again.

Test this solution for **K**.

Where does the **K** in the ashes come from? Give its history.

Add a few drops of **HCl** to the solution. What gas is given off?

What salt of **K** is in the solution?

WEIGHTS AND MEASURES.

APOTHECARIES' MEASURE.

60 minims	make	1 fluid drachm.
8 fluid drachms	"	1 fluid ounce.
16 fluid ounces	"	1 pint.
8 pints	"	1 gallon.

APOTHECARIES' WEIGHT.

20 grains	make	1 scruple.
3 scruples	"	1 drachm.
8 drachms	"	1 ounce.
12 ounces	"	1 pound.

COMPARATIVE, FLUID MEASURE.

1 c. c. =	17	minims (nearly).								
4 c. c. =	63	"	= 1	drachm	and	8	minims.			
5 c. c. =	85	"	= 1	"	"	25	"			
8 c. c. =	136	"	= 2	"	"	16	"			
10 c. c. =	170	"	= 2	"	"	50	"			
30 c. c. =	510	"	= 1	ounce	"	0	drachm	30	minims	
50 c. c. =	850	"	= 1	"	"	6	"	10	"	
60 c. c. =	1020	"	= 2	"	"	1	"	0	"	
100 c. c. =	1700	"	= 3	"	"	4	"	20	"	

COMPARATIVE, SOLID MEASURE.

1 gram	=	$15\frac{2}{3}$	grains.				
5 "	=	77	"	= 1	drachm and	17	grains
8 "	=	$123\frac{1}{5}$	"	= 2	" "	$3\frac{1}{5}$	"
10 "	=	154	"	= 2	" "	34	"
20 "	=	308	"	= 5	" "	8	"
50 "	=	770	"	= 12	" "	50	"
100 "	=	1540	"	= 25	" "	40	"

METRIC SYSTEM.

LINEAR MEASURE.

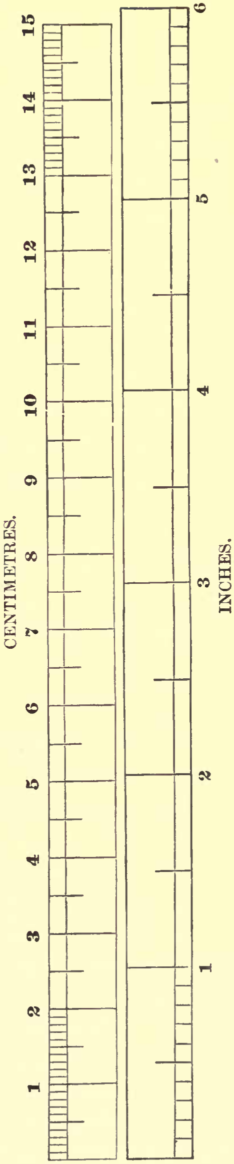
10 millimetres (mm.)	= 1 centimetre (cm. or c.)	=	.39 in.
10 centimetres	= 1 decimetre (dcm.)	=	3.93 "
10 decimetres	= 1 metre (m.)	=	39.37 "
10 metres	= 1 dekametre.		
10 dekametres	= 1 hektometre.		
10 hektometres	= 1 kilometre.		

CAPACITY.

10 millilitres	=	1 centilitre.
10 centilitres	=	1 decilitre.
10 decilitres	=	1 **litre** = 61 cu. in.
10 litres	=	1 dekalitre.
10 dekalitres	=	1 hektolitre.
10 hektolitres	=	1 kilolitre.

WEIGHT.

10 milligrams (mg.)	=	1 centigram (cgm.).
10 centigrams	=	1 decigram (dcg.).
10 decigrams	=	1 **gram** = 15.43.
10 grams	=	1 dekagram.
10 dekagrams	=	1 hektogram.
10 hektograms	=	1 kilogram (kgm.) = $2\frac{1}{5}$ lbs.



Do not try to translate these weights and measures into common or U. S. terms. One can very quickly learn the three units, the Gram, the Litre, and the Metre, and then the important divisions and multiples of these units. Examine and use the weights and measures until they become as familiar as ounces, pints, or pounds. In writing the fractions of a unit use the decimal system, thus:

Millimetre,	0.001
Centimetre,	.01
Decimetre,	.1
Metre,	1.

A cubic centimetre of distilled water at 4° C. weighs *one gram.* One litre = a cubic decimetre, or a thousand cubic centimetres = 1.05672 U. S. quarts.

THERMOMETRIC RULES.

To change Fahrenheit degrees to Centigrade, subtract 32° and multiply the remainder by $\frac{5}{9}$. C. = (F. — 32) $\frac{5}{9}$.

To change Centigrade degrees to Fahrenheit, multiply by $\frac{9}{5}$ and add 32°. F. = $\frac{9}{5}$ C. + 32°.

STANDARD SOLUTIONS.

Iodine Solution for Starch Test.

To 1 gram of metallic iodine and 1.5 grams of dry potassium iodide add 100 c. c. of distilled water. Stir with a glass rod until all the iodine dissolves. The object of the potassium iodide is to increase the solubility of the iodine. This solution will not keep well, and so should be freshly made for the tests upon digestion of starch.

Copper Sulphate Solution for Detection of Grape Sugar—Fehling's Fluid.

(*a*) Dissolve 3.46 grams of crystallized cupric sulphate in 16 c. c. of distilled water.

(*b*) Dissolve 17.3 grams of pure crystallized potassium-sodium tartrate in 60 c. c. of sodium hydrate, specific gravity 1.12.

Add (*a*) to (*b*), stirring well; then dilute to 500 c. c. with distilled water. This fluid will not keep long, so it is best to prepare it each time it is required. If kept cool and dark in full, glass-stoppered bottles, it will keep for a week or more. It is well to test it by boiling, to see if it will reduce alone, before using it for sugar test. .005 gram of grape-sugar will completely reduce 1 c. c. of this solution.

REAGENTS.

Strength of Solutions Usually Employed.

Acids:

Acetic (vinegar), 30%, specific gravity 1.04.

Hydrochloric (Muriatic), 24%, specific gravity 1.12.

Hydrosulphuric, gas or saturated solution.

Nitric, 32%, specific gravity 1.2.

Sulphuric (Oil of Vitriol), concentrated 1.8. Dil. 1 to 5 water.

Aqua regia, 1 part concentrated $\mathbf{HNO_3}$ to 4 parts **HCl**.

Ammonium:

" Chloride, 1 part crys. salt to 8 water.

" Carbonate, 1 part crys. salt to 4 water + 1 part $\mathbf{NH_4HO}$.

" Hydrate, 12%, specific gravity .95.

" Sulphide, $\mathbf{NH_4HO}$ saturated with $\mathbf{H_2S}$.

Barium Chloride, 1 part to 10 of water.

" Hydrate, 1 part to 20 in $\mathbf{H_2O}$.

Calcium " saturated solution in $\mathbf{H_2O}$.

" Sulphate, " " "

Calcium Chloride,	1	part	to	8	**H_2O**.
Cobalt Nitrate,	1	"	"	8	"
Copper Sulphate,	1	"	"	8	"
Ferrous Sulphate,	1	"	"	5	"
Ferric Chloride,	1	"	"	15	"
Lead Acetate,	1	"	"	10	"
Magnesium Sulphate,	1	"	"	10	"
Mercuric Chloride,	1	"	"	16	"
Potassium Bichromate,	1	"	"	10	"
" Cyanide,	1	"	"	4	"
" Ferricyanide,	1	"	"	12	"
" Ferrocyanide,	1	"	"	12	"
" Iodide,	1	"	"	20	"
" Sulphate,	1	"	"	12	"
Sodium Carbonate,	1	"	"	5	"
" Hydrate,	1	"	"	20	"
" Phosphate,	1	"	"	10	"
Silver Nitrate,	1	"	"	20	"
Stannous Chloride,	1	"	"	6	" + **HCl** (acidulate).

A KEY FOR DETERMINING THE BASES OF UNKNOWN COMPOUNDS.

A. Substances soluble in water or in acids.

To the solution add **HCl.**

(*a*) A precipitate shows.

1. **AgCl.** Soluble in **NH_4HO.**
2. **$HgCl_2$.** Turns black with **NH_4HO.**
3. **$PbCl_2$.** Soluble in hot water, not in **NH_4HO.**

(*b*) No ppt. Add **H_2S** to the acid solution.

* A ppt. Neutralize with **NH_4HO,** and add **$(NH_4)_2S$** in excess, and

† The ppt. dissolves.

1. **As_2S_3.** The ppt. obtained with **H_2S** was yellow.
2. **Sb_2S_3.** " " " " " " orange.
3. **SnS_2.** " " " " " " yellow.

Put some grains of zinc into a test-tube and cover with dilute **H_2SO_4.** Add to this some of the original solution. Cover the test-tube with a filter-paper moistened with **$HgCl_2$.** Arsenic stains the paper *yellow* or orange; antimony, gray or black; tin, no change.

4. **SnS.** " " " " " " brown.

†† The ppt. does not dissolve in **$(NH_4)_2S$.**

1. **HgS.** Ppt. is black. Put original solution upon clean copper coin. It deposits mercury. Rub the coin with a towel.
2. **PbS.** " " " Original solution gives chrome-yellow ppt. with **$K_2Cr_2O_7$.**
3. **BiS.** " " " Add water to original solution and it becomes milky.
4. **CuS.** " " " **NH_4HO** and original slution give a deep blue solution.
5. **CdS.** " " yellow. The yellow sulphide is characteristic.

** No ppt. Add **NH_4Cl** to the same, and then **NH_4HO,** until it smells of ammonia, then add **$(NH_4)_2S$.**

‡ A ppt. which is

1. Black.

(*a*) **FeS.** Test original solution with **K_3Cfy.** Blue = ferrous; greenish = ferric.

(*b*) **CoS.** Blue borax bead.

(*c*) **NiS.** Brown borax bead.

2. Light.

(*a*) $\mathbf{Al_2H_6O_6}$. White ppt. Blue mass on charcoal with $\mathbf{Co(NO_3)_2}$.
(*b*) **ZnS.** " " Green " " " " "
(*c*) $\mathbf{Cr_2H_6O_6}$. Green " Green borax bead.
(*d*) **MnS.** Light-gray ppt. Green mass on platinum-foil when fused with $\mathbf{Na_2CO_3}$ and $\mathbf{KNO_3}$. Amethyst borax bead.

‡‡ No ppt. Boil to expel $\mathbf{H_2S}$, or to the original solution add $\mathbf{NH_4HO}$, $\mathbf{NH_4Cl}$, and $\mathbf{(NH_4)_2CO_3}$.

I. A ppt. indicates:

1. $\mathbf{BaCO_3}$. Burn some of the dry salt on the platinum wire, and it colors the flame yellowish green.
2. $\mathbf{SrCO_3}$. " " " " " " " " " " crimson.
3. $\mathbf{CaCO_3}$. " " " " " " " " " " yellow to orange.

II. No ppt. Add to the same solution $\mathbf{Na_2HPO_4}$.

(*a*) A ppt.

$\mathbf{MgCO_3}$.

(*b*) No ppt.

1. $\mathbf{NH_4}$. Add **KHO** to the solution, or rub slaked lime with the original salt in the mortar, and it gives off ammonia, $\mathbf{NH_3}$.
2. **Na.** Burn on platinum wire and it colors the flame intensely yellow.
3. **K.** Burn on platinum wire and it colors the flame violet. Look through blue glass to shut off sodium, which is apt to obscure test.

B. Insoluble in acids.

Pulverize, if possible, and mix with four times its weight of $\mathbf{Na_2CO_3}$. Put into a porcelain crucible, and heat before the oxhydrogen blow-pipe, gently at first, then strongly, until the mass is fused. When cool, put the crucible and contents into a beaker of water, and heat till the mass dissolves. Now filter and wash. The filter will contain the base, and the liquid, the acid. Dissolve the substance in the filter in $\mathbf{HNO_3}$, and proceed according to **A.** Save the liquid for the next.

A KEY FOR DETERMINING THE ACID OF AN UNKNOWN SALT.

A. Chars or swells and turns black when heated upon a piece of porcelain.

1. Tartaric acid, $H_2C_4H_4O_6$. Gives the odor of burnt sugar while burning.
2. Oxalic " $H_2C_2O_4$. Add to solution MnO_2 and H_2SO_4, and it will give off CO_2. Incline test-tube over a test-tube of lime-water, cover, and shake.
3. Acetic " $HC_2H_3O_2$. Add H_2SO_4 and an equal volume of alcohol, heat, and an agreeable odor of fruit is given off.

B. Does not char nor turn black. Add dil. HCl to a solution of the salt.

(*a*) A gas is evolved. This shows one of the following acids:

1. H_2CO_3. Incline the test-tube over a test-tube of lime-water, so that the gas will drop into the lime-water tube. Cover the latter and shake. It becomes clouded.
2. H_2S. Cover the mouth of test-tube with filter-paper moistened with lead acetate. It will be stained black. Also odor of rotten eggs.
3. H_2SO_3. Gives odor of burning sulphur.
4. $HClO$. Bleaches moistened litmus paper.
5. HNO_2. Add acetic acid and $FeSO_4$, and a dark-brown color is produced.

(*b*) No gas is evolved. Add to the original solution, $BaCl_2$.

* A ppt. is formed, insoluble in aqua regia.

H_2SO_4. Sufficiently proven.

* No ppt. is formed. Add HNO_3, $CaSO_4$, and then NH_4HO in excess to the original solution.

† A ppt. is formed.

(*a*) The ppt. dissolves upon the addition of acetic acid.

1. Phosphoric acid, H_3PO_4. Add the original solution drop by drop to a solution of $(NH_4)_2MoO_4$ in HNO_3, and warm. A yellow ppt. forms.
2. Silicic acid, H_2SiO_3. Add HCl in excess and evaporate to *dryness* on the water-bath. A white powder, SiO_2, remains insoluble in acids.

(*b*) The ppt. does not dissolve in acetic acid.

1. Oxalic acid. (See above.)
2. Hydrofluoric acid. Pulverize original substance, put into a test-tube, and add H_2SO_4. Insert a smaller empty test-tube and the gas evolved will etch it near the top.

†† No ppt. is formed with acetic acid. To original solution add $\mathbf{HNO_3}$ and $\mathbf{AgNO_3}$.

x. A ppt. is formed which is

1. Soluble in $\mathbf{NH_4HO}$.

(*a*) **HCl.** Add a few drops of concentrated $\mathbf{H_2SO_4}$, and then $\mathbf{MnO_2}$, and chlorine will be evolved.
(*b*) **HBr.** Add a few drops of $\mathbf{HNO_3}$ to original solution, and then some $\mathbf{H_2SO_4}$, and heat. Red-brown liquid.
(*c*) **HCy.** Add to some of the original solution in a porcelain dish a small quantity of $\mathbf{(NH_4)_2S}$, and digest on the water-bath until all the sulphide is destroyed. Then acidulate with **HCl** and add $\mathbf{Fe_2Cl_6}$, and a blood-red solution will be formed.
(*d*) $\mathbf{H_3Cfy}$. Add $\mathbf{FeSO_4}$ to original solution, and a Prussian-blue ppt. forms.

2. Insoluble in $\mathbf{NH_4HO}$.

(*a*) **HI.** Add $\mathbf{MnO_2}$ and $\mathbf{H_2SO_4}$ to original solution, and the gas evolved will stain starch blue.
(*b*) $\mathbf{H_4Cfy}$. Add $\mathbf{Fe_2Cl_6}$ and a dark-blue ppt. forms.

y. No ppt. is formed with $\mathbf{AgNO_3}$. Heat a small portion of the dry substance on charcoal before the blow-pipe.

1. Burns suddenly or deflagrates.

(*a*) $\mathbf{HNO_3}$. To the solution add an equal volume of concentrated $\mathbf{H_2SO_4}$, and when cold pour down the side of the tube a solution of $\mathbf{FeSO_4}$. A brown ring forms.
(*b*) $\mathbf{HClO_3}$. Heat the dry solid in a hard glass tube, and oxygen will be liberated.

2. Does not burn suddenly.

(*a*) $\mathbf{HBO_2}$, Boric acid. Put into a porcelain dish. Add concentrated $\mathbf{H_2SO_4}$, then alcohol, and ignite. Burns with a green flame.
(*b*) $\mathbf{H_2CrO_4}$, Chromic acid. Gives chrome yellow, with lead acetate.
(*c*) $\mathbf{HC_2H_3O_2}$. (See acetic acid above.)

C. If no acid is found it may be an oxide, a hydrate, or an elementary substance, and must be looked for in the analysis of *bases*.

QUESTIONS SUGGESTING FURTHER STUDY OF THE METALS.

Of what use is mercury in gold mining? Dissolve gold-foil in a globule of mercury. What is an amalgam? What per cent. of gold in gold coins? $\mathbf{FeSO_4}$ will precipitate gold from its solutions, and aqua regia will dissolve it, forming a chloride. $\mathbf{3FeSO_4 + AuCl_3 = Fe_2(SO_4)_3 + FeCl_3 + Au}$ and $\mathbf{Au + 3Cl = AuCl_3}$. What is brass? How could you separate it into its constituents? What is plumbers' solder? What is the probability that any of the metallic elements may be found to be compounds? (See Professor William Crookes' article in *Forum* for May, 1891.) Why are bicarbonate of sodium and potassium useful in cake-making? Why is yeast used in making bread? Could bread be made light without any of these? How? How does plaster of Paris differ from gypsum? How is quick-lime obtained? What metals are useful in medicine? If potassium, sodium, and ammonia are so similar in all their uses, is there any probability of the resolution of the two former into other elements, as ammonia is? Has any prominent scientist ever announced a belief in all elements being some time reduced to a form of hydrogen? What metals are found free in nature? How do other metals, when in an elementary state, even the common ones, compare in value with gold? Why is gold called a precious metal? What metals were first known to the ancients? What uses could the metal aluminum be put to when cheap enough for general use?

To the Teacher.—The pupil should now be given a number of substances to determine first the base, then the acid. These should be combined and the name and symbol of the salt written out. It will be noticed that a different method, to some extent, has been used in forming a Key, to that used in the experiments upon metals. The Key is only a brief method, and offers some advantages in rapid work, but the pupil should confirm the results obtained here by referring to the experiments in the text for further proof of his work.